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HEAT

by

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"Heat itself is Motion and nothing else"

FRANCIS BACON

CAMBRIDGE
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PREFACE

This book, following my Hydrostatics and Mechanics, marks the second stage in a course of Physics of School Certificate standard.

The pupils for whom it is designed are, as Sir Percy Nunn would say, midway between the "wonder" and "utility" stages of their development: not till they begin to specialise do they reach the "system" stage. Their interest, therefore, in the study of heat lies primarily in the phenomena and their own experiments, but also in such machines as the refrigerator, the steam and internal combustion engines.

I have endeavoured, without any sacrifice of rigid scientific treatment, to make the book readable, attractive and human. There are numerous illustrations and descriptions of the practical applications of the principles, besides summaries and School Certificate questions at the end of each chapter.

There is a small point of interest which may be mentioned here. I have omitted Davy's well-known "ice rubbing" experiment in view of Prof. Andrade's criticism in *Nature* (March 9th 1935).

I am very much indebted to Mr D. G. A. Dyson of King Edward VI School, Stratford-on-Avon, for reading the manuscript and making numerous criticisms and suggestions; for taking the photographs of the two-stroke engine, and a cloud forming over a mountain. I am also indebted to my colleague, Dr A. W. Barton, for reading the manuscript and making valuable suggestions. Mr C. J. P. Cave has kindly given me permission to use some of his pictures of clouds. I am most grateful to my pupils Mr A. Lyell and Mr W. D. M. Paton for working out the answers to the examples.

Messrs The British Thomson-Houston Co., C. A. Parsons & Co. Ltd., the Cambridge Instrument Co. Ltd., Newalls Insulation Co. Ltd., and Babcock and Wilcox, have all kindly lent me

literature and photographs; both these and other acknowledgments have been made under the illustrations.

Finally I wish to thank the Oxford and Cambridge Joint Board, the Northern University Joint Matriculation Board, the Cambridge Local and the Oxford Local Examinations Syndicates who have given me permission to reproduce School Certificate Questions.

A. E. E. M.

Repton
October 1935



HEAT

Chapter I

TEMPERATURE AND THERMOMETERS

The study of heat began with the invention of instruments for measuring hotness.

Water in a kettle, when placed on a fire, becomes hotter. Its degree of hotness (at any instant) is called its *temperature*, and an instrument for measuring it is known as a *thermometer*.

For many thousands of years men were content to rely on their sensations of warmth and cold for gauging the hotness of a body. These sensations, of course, give a vague, inaccurate, and sometimes misleading estimate of temperature. The scientific study of heat began when attempts were made to obtain accurate measurement of temperature.

The first thermometer.

When bodies are heated they expand. The first thermometer, invented by Galileo about the year 1593, depends on this fact.

It is represented in Fig. 1 (a). It consists of a glass flask full of air, turned upside down, and fitted with a cork through which a glass tube passes. The glass tube dips into water, and the quantity of air in the flask is arranged so that the water rises a convenient distance up the tube.

When the flask is warmed, the air in it expands, and pushes the water down. A scale is placed behind the tube, and the movement of the water up and down registers changes of temperature.

This thermometer is subject to two obvious errors. The evaporation of the water tends to alter its readings (would this make the readings higher or lower?), and fluctuations in the atmospheric pressure cause

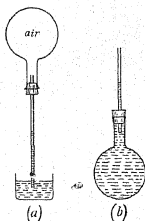


Fig. 1

the water to move up and down even though the temperature remains constant.

It is called an air thermometer, because its working depends on the expansion of air when heated. To-day air thermometers are still made, similar in principle to Galileo's, but modified to eliminate errors.

Liquid thermometers.

The next type of thermometer to be invented depended not upon the expansion of air, but upon the expansion of a liquid when heated. The flask, in this case not inverted (see Fig. 1(b)), is fitted with a glass tube, and is filled with a liquid such as water. When the temperature of the liquid rises, its level ascends up the tube, behind which a scale is placed. This instrument has the advantage that variations in atmospheric pressure have no effect upon its readings.

The advantage of sealing the top of the tube to prevent evaporation soon became evident, and the relative sizes of the flask and tube were adjusted so that the difference between summer and winter temperatures was represented by the liquid level moving the whole length of the tube. Alcohol was found to be a better liquid than water as it did not freeze solid in the winter and crack the flask.

Later an instrument maker named Fahrenheit filled his thermometers with mercury instead of alcohol. Many people preferred alcohol, as it expands far more than mercury. Mercury, however, has two great advantages over alcohol which finally decided in its favour (see p. 5).

The next improvement in the thermometer was a decrease in size, making it much more compact and easier to use. But if the size of the flask is decreased, the bore of the tube must be correspondingly decreased. Hence a glass tube with a very fine bore known as a capillary tube (*capillus*—a hair) is always used.

How to make a mercury thermometer.

Heat the end of a capillary tube in a blow pipe flame and, slowly turning it, seal it up. Blow down it while it is still hot, and produce a bulb on the end. A bulb is far neater than a flask, and can be blown to a convenient size according to the length of tube we wish to use.

The next problem is to fill the thermometer with mercury. Fix a funnel to the top by means of a short piece of rubber tubing, and pour mercury into the funnel. The mercury will not pour into the tube as the latter is too narrow for the air to escape while the mercury runs down.

To remove the air heat the bulb gently with a bunsen flame (see Fig. 2). This causes the air to expand and some of it pushes its way out, bubbling up through the mercury. On cooling the bulb, the air contracts, and some of the mercury enters the bulb. By alternately heating and cooling in this manner the bulb can be filled. To drive out the last vestiges of air the mercury must be boiled.

When the bulb and tube are full to the top, we have to decide exactly how much mercury we require. Suppose we wish the thermometer to read up to the temperature of boiling water: then the mercury at this temperature should be just a few millimetres from the top of the tube.

Heat the bulb in boiling water. Mercury will be driven out until the level at this temperature is exactly at the top of the tube. Heat the bulb to a slightly higher temperature over a flame, and so drive out a little more mercury. Allow the bulb to cool slightly, then, while the mercury is near the top of the tube, seal it with a blow pipe flame to prevent air from entering.

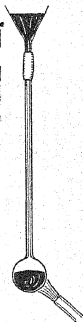


Fig. 2

Graduating the thermometer.

How are we to graduate the thermometer now it is made? Suppose we put 0 at the bottom of the stem, 10 or 100 at the top, and divide the space between into equal divisions. Will all thermometers made in this way read the same? The volumes of the bulbs are bound to vary slightly (to make them all the same size is impracticable), the stems may not be exactly the same length, and no precaution has been taken to see that they are filled with the same volume of mercury. Thus all thermometers so graduated will read differently.

The first makers were faced with this problem. Where were they to start marking divisions on the stem, and exactly how wide were they to make them in each case?

They perceived that the solution to the problem was to mark the position of the mercury in each stem at two fixed temperatures, called the *fixed points*, and divide the intervening space into the same number of divisions. Although these divisions in different thermometers differ in length, they represent the same change of temperature.

During the seventeenth century, the best thermometers were made in Florence in Italy. The Florentines took as their lower fixed point the temperature of a deep cellar, which is that of the earth at a certain depth, and remains fairly constant all the year round; and as their upper fixed point, the temperature of the bodies of healthy animals, cows and deer, which is also a fairly constant temperature.

Later, better fixed points were discovered. It was found that pure melting ice is always at the same temperature; and that water boils at a certain steady temperature, extra heat having no effect in raising its temperature but merely causing it to boil away more quickly. These fixed points are now used for all thermometers.

Thus to determine the lower fixed point of our thermometer place it in melting ice preferably in a funnel to allow the water to drip away, leave it for some minutes, and inspect it at intervals to see if the level of the mercury has ceased to drop. When the level is quite steady, mark its position by a rubber band or a scratch on the glass with a file.

Why must the ice be melting? Is, for instance, the temperature of the ice at the North Pole the same as that of melting ice?

When salt is mixed with ice the temperature drops; hence the ice must be absolutely pure.

To mark the upper fixed point, place the thermometer in pure boiling water, and make another scratch. We shall see later that a more careful determination of the upper fixed point must be made, however, if the thermometer is to be really accurate.

Scales of temperature.

When the fixed points are marked, we have to decide what numbers we shall assign to them. The simplest scale of temperature is the Centigrade scale, on which the lower fixed point is marked 0°C. , and the upper fixed point is marked 100°C. This scale is used in France and, because of its simplicity, by

scientists all over the world. It is interesting to note that the inventor, Celsius, called the melting point of ice 100, and the boiling point of water 0. There is, of course, no real objection to this.

An earlier scale, and the one which is chiefly in use in England and America, is the Fahrenheit scale. The lower fixed point is marked 32°F ., and the upper, 212°F .

The origin of the numbers 32 and 212 is a somewhat involved story. Fahrenheit wanted to take 0° as the lowest temperature he could reach (he obtained it by mixing ice and salt) and 96° as the temperature of the human body. To do this he had to make his fixed points 32° and 212° . But it was found later that the accurate temperature of the human body on this scale was 98.4° .

There is another scale of temperature, which is widely used in Germany and Russia, the Réaumur scale. The fixed points are 0°R . and 80°R .

The thermometers on some old barometers are marked with a forgotten scale known as the Royal Institution scale. This is only one of at least a score which were invented in the eighteenth century. "All but three of them have passed into oblivion. Would that the Centigrade scale were the sole survivor!" (Cajori)—a feeling which every schoolboy will share.

Why mercury is used.

The chief advantage of mercury for use in a thermometer is its large range. It does not freeze solid until -39°C ., a lower temperature than is ever reached in this country, and does not boil until 357°C . Alcohol (ethyl), its chief rival, on the other hand, boils at 78°C . However, alcohol has a much lower freezing point than mercury, -112°C ., and is convenient for use on polar expeditions. Other advantages of mercury are that it very rapidly reaches the temperature it is required to measure, while alcohol takes some time; it is easily seen; it does not, like alcohol, wet the sides of the tube and hence does not tend to read too low when the temperature is falling rapidly.

More accurate determination of the upper fixed point.

The addition of salt or similar impurity raises the temperature of boiling water. A concentrated salt solution, for example, boils at 107°C . If however, a thermometer is placed in the steam from

the boiling solution it records 100°C . Faraday's announcement of this discovery in 1822 was met with some scepticism. Gay-Lussac expressed his conviction that the steam as it leaves the solution must be at the same temperature as the solution. Some of the steam, however, condenses on the bulb of the thermometer, which therefore records the boiling point of pure water. The temperature of the condensed steam cannot rise above its boiling point: when heated it merely tends to boil away.

Hence a thermometer which is being graduated is always placed in the steam from boiling water. This eliminates any error due to impurities in the water.

Again, the boiling point of water rises slightly if the atmospheric pressure increases. Under the very large pressure (about 16 atmospheres) in the boiler of a modern locomotive, water boils at 200°C . *The upper fixed point is defined as the temperature of the steam from water boiling under a pressure of 76 cm. of mercury.* Hence the barometer must be read and a correction made when determining the upper fixed point.

A special boiler (see Fig. 3) called a hypsometer is used for testing the upper fixed point of a thermometer. The bulb and most of the stem of the thermometer are immersed in the steam from boiling water. The vessel is double-walled so that the inner steam chamber is surrounded by an outer one to keep it hot. A small water manometer indicates whether the pressure inside the apparatus is the same as that of the outside atmosphere. The two levels in the manometer should be the same, and if not, a correction applied to the measured atmospheric pressure.

The word "hypsometer" means "height-measurer" for it is possible to calculate the height of a mountain from the temperature at which water boils at the top.

At the summit of Mont Blanc (15,000 ft.) water boils at about 84°C . Thus it is impossible to make good tea or cook an egg at such heights without using a pressure cooker in which the steam is allowed to increase the pressure in the apparatus considerably above that of the atmosphere around, thereby raising the boiling point of water to its normal value.

Pressure cookers are also used at ordinary altitudes. The

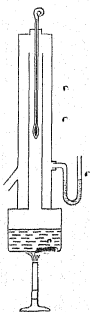


Fig. 3. Hypsometer.

cooker is constructed so that steam accumulates inside it until at a certain pressure it blows off through a valve. The water is made to boil at a temperature of about 110°C. , enabling a chicken to be cooked in less than half an hour.

The National Physical Laboratory.

Thermometers, however accurately they may have been graduated, are apt to develop errors owing to the gradual shrinkage of the glass bulb with age. They may be tested at the National Physical Laboratory, Teddington, for a small charge. The N.P.L., as it is often called, was founded in 1900 with a staff of less than 30. Its staff, now more than 600 strong, besides testing standards, are occupied in all kinds of physical research, in aeronautics, ship design, wireless, metallurgy, etc.

Conversion of scales.

To convert 15°C. to $^{\circ}\text{F.}$ Between the two fixed points there are 100 Centigrade degrees. Between the two fixed points there are $212 - 32 = 180$ Fahrenheit degrees.

$$\begin{aligned} \therefore 100 \text{ C. degrees} &\equiv 180 \text{ F. degrees} \\ 1 \text{ C. degree} &\equiv \frac{180}{100} \text{ F. degrees} \\ &\equiv \frac{9}{5} \text{ F. degrees} \end{aligned}$$

$$\text{Thus } 15 \text{ C. degrees} \equiv 15 \times \frac{9}{5} = 27 \text{ F. degrees.}$$

But the temperature 15°C. cannot be 27°F. , for the latter is actually below freezing point. Since the freezing point on the Fahrenheit scale is not 0°F. but 32°F. we must add 32 to our value (see Fig. 4).

$$\begin{aligned} \text{The temperature } 15^{\circ}\text{C.} &= 27 + 32 \\ &= 59^{\circ}\text{F.} \end{aligned}$$

To convert 95°F. to $^{\circ}\text{C.}$

$$\begin{aligned} 180 \text{ F. degrees} &\equiv 100 \text{ C. degrees.} \\ 1 \text{ F. degree} &\equiv \frac{5}{9} \text{ C. degree.} \end{aligned}$$

Now a temperature of 95°F. is $(95 - 32) = 63$ degrees above freezing point.

$$63 \text{ F. degrees} \equiv 63 \times \frac{5}{9} \equiv 35 \text{ C. degrees.}$$

$$\therefore \text{The temperature } 95^{\circ}\text{F.} = \text{the temperature } 35^{\circ}\text{C.}$$

The reader should now be able to see for himself how to carry out conversions to the Réaumur scale.

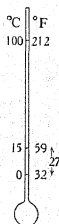


Fig. 4

The clinical thermometer.

The clinical thermometer is a special type of thermometer used by doctors for finding the temperature of their patients. The normal temperature of the human body is 98.4°F. , and cannot vary by many degrees without death resulting. Hence the clinical thermometer is graduated only between 95°F. and 110°F. It has a small kink or constriction in the bore of the stem (see Fig. 5), which prevents the mercury from running back into the bulb when the thermometer is taken out of the patient's mouth and read by the observer at room temperature. The mercury is made to run back through the constriction, before a fresh reading

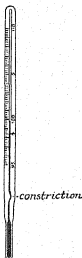


Fig. 5



Fig. 6



Fig. 7

is taken, by vigorous shaking of the thermometer. It has no difficulty in forcing its way up through the constriction since it is capable of exerting great force during expansion.

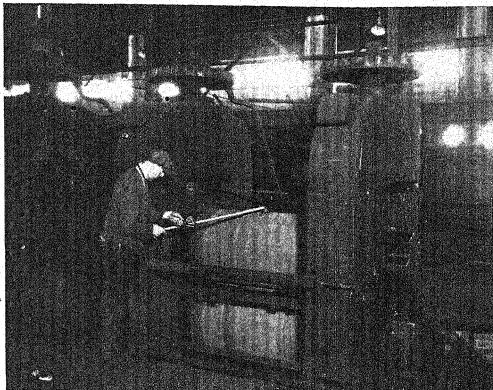
Maximum and minimum thermometers.

It is sometimes necessary to know the maximum and minimum temperatures of a room, such as a greenhouse, during a period of 24 hours. Special self-recording thermometers have been invented for this purpose.

The usual principle on which they work is the forcing of a small index by the meniscus of the liquid in the thermometer to the maximum or minimum position, where it remains when the meniscus recedes. In a maximum thermometer mercury, which

has a convex meniscus, is used; in a minimum thermometer, alcohol, which has a concave meniscus (see Fig. 6).

Six's combined maximum and minimum thermometer is much used. It consists (see Fig. 7) of a bulb A completely filled with alcohol, a U-tube B containing a long thread of mercury, a bulb C partially filled with alcohol, and two indices, I_1 and I_2 .



By courtesy of the Cambridge Instrument Co. Ltd.

Fig. 8. Taking the temperature, by means of a thermo-electric pyrometer, of the hot rollers used for rolling sheet steel. If steel is passed between the rollers when they are cold considerable damage may be done to them. Note the galvanometer in the man's left hand: it is graduated direct in $^{\circ}\text{C}$.

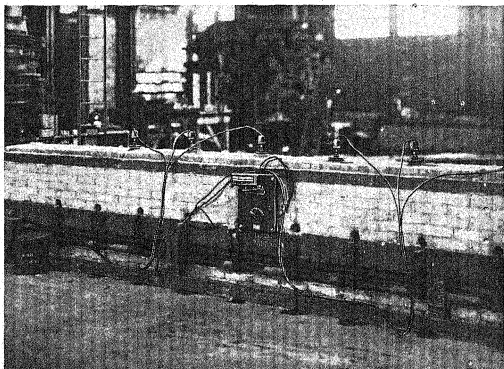
When the temperature rises the alcohol in A expands with the result that the index I_1 is forced up by the mercury meniscus. Thus I_1 records the maximum temperature.

When the temperature falls the alcohol in A contracts, and owing to the pressure of the alcohol vapour and air in C , the index I_2 is forced up. Thus I_2 records the minimum temperature. The instrument requires special calibration.

Another way of measuring temperature. The thermo-couple.

There is a method of measuring temperature entirely different from that used in the mercury and gas thermometers.

Wires of two different metals, say iron and copper, are joined, and their other ends connected to a sensitive instrument for detecting electric current, called a galvanometer. When the junction



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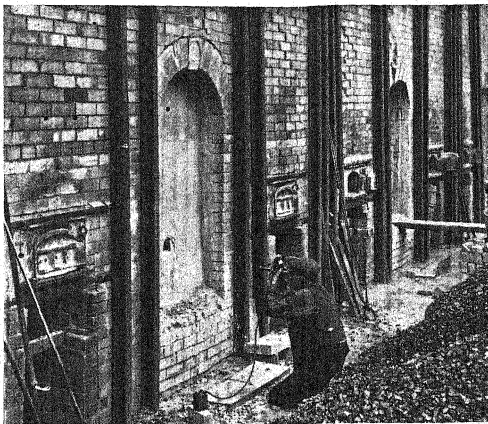
Fig. 9. Six thermo-electric pyrometers used to record the temperature at different parts of a glass lehr, i.e. a gas-heated furnace in which the glass is annealed, a process of cooling at a definite rate to avoid strain in the glass. The recording instrument is shown in the centre of the picture, and it may be connected to any of the pyrometers by means of the switch. This is a compact illustration of distance thermometry.

of the metals is heated by a match, the pointer of the galvanometer is deflected, showing that an electric current is passing. The size of the electric current (of the order of a few milli-amperes) depends on the temperature of the junction, and can therefore be used to measure it. The phenomenon is called the thermo-electric effect and two wires used in this way are known as a thermo-couple.

Pyrometers.

An instrument for measuring high temperatures is called a pyrometer.

The majority of those used in industry—for measuring the temperatures of furnaces, molten metals, the steam or exhaust



By courtesy of the Cambridge Instrument Co. Ltd.

Fig. 10. Taking the temperature of a brick kiln with a disappearing filament radiation pyrometer. The workman matches the light of the furnace, viewed through a hole in the side, with the light of an electric lamp filament whose brightness can be altered by varying the current. The filament is arranged to overlap the image of the hole in the furnace. When the current is too small or too large, the filament seems darker or brighter, respectively, than the hole: when the adjustment is correct, the filament disappears.

gases of engines, the rollers of a steel rolling mill, etc.—are of the thermo-electric type (see Figs. 8 and 9). An advantage of these pyrometers is that they may be connected by wires to galvano-

meters in a central control room where the temperatures of distant or inaccessible parts of an industrial plant may be read.

The thermo-electric pyrometer will not stand up to temperatures higher than 1250°C . For temperatures between 1250°C . and 3000°C . a radiation pyrometer is used (see Fig. 10).

Extreme temperatures.

What are the lowest and highest temperatures that can be reached? We have reason to believe that there is no such thing as a temperature lower than -273°C . and this is known as Absolute Zero. Temperatures have been reached actually within a small fraction of a degree of Absolute Zero in the laboratory of Prof. Onnes at Leyden.

Air becomes liquid at about -180°C . and substances subjected to this intense degree of cold undergo remarkable changes. A grape or rubber ball, for instance, becomes so brittle that it can be broken into fragments with a hammer.

There is no theoretical upper limit to temperature. The highest (enduring) temperature obtained by man is about 3500°C ., that of the electric arc. (Temperatures of about $20,000^{\circ}\text{C}$. have been achieved momentarily by "exploding" fine wires by very large electric voltages.) The temperatures of some of the stars, however, make this look small by comparison. Their surface may be at $20,000^{\circ}\text{C}$. and their interior, $30,000,000^{\circ}\text{C}$. or higher.

EXPERIMENT 1. *To compare a Centigrade and Fahrenheit thermometer.*

Place a Centigrade and Fahrenheit thermometer in a can of water. Heat the water slowly and read the thermometers simultaneously at intervals of about 10°C ., stirring steadily.

Plot $^{\circ}\text{C}$. against $^{\circ}\text{F}$. Use the graph to convert various temperatures from one scale to the other and check by calculation.

EXPERIMENT 2. *An experiment on the lower fixed point of a thermometer.*

Completely surround the bulb of a thermometer with melting ice. Add salt to the ice and note the effect on the thermometer.

EXPERIMENT 3. *An experiment on the upper fixed point of a thermometer.*

The upper fixed point is the temperature of steam from water boiling at standard atmospheric pressure.

The object of the experiment is to show

1. The effect of change of pressure,
 2. The effect of impurities (e.g. salt),
- on the temperature of boiling water, and the steam from it.

Use the apparatus shown in Fig. 11 containing distilled water.

1. Boil the water and take the temperature of the steam from it.
2. Increase the pressure of the steam by holding a ruler over *A* and record the maximum temperature reached.

3. Decrease the pressure inside the flask by taking away the bunsen flame and clipping the rubber tube at *A*. The steam inside the flask condenses leaving a partial vacuum. Record the lowest temperature at which the water boils. The flask may be cooled under the cold water tap to accelerate the condensation of the steam.

4. Lower the thermometer into the water and find the temperature of the boiling water.

5. Draw up the thermometer again. Add salt to the water and find the temperature of the steam taking care to prevent salt solution splashing on to the bulb of the thermometer.

6. Lower the thermometer and find the temperature of the boiling salt solution.

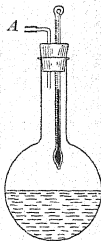


Fig. 11

EXPERIMENT 4. *To make and graduate a thermometer.*

It is not difficult to blow a bulb at the end of a glass capillary tube and to make and graduate a thermometer for oneself. Follow the instructions on pp. 2-4. A cheap and convenient liquid to use in the thermometer is amyl alcohol.

SUMMARY

Temperature is the degree of hotness of a body. The commonest type of thermometer utilises the expansion of a liquid to measure temperature. Such a thermometer is graduated by marking on it two fixed points: (1) the temperature of pure melting ice, (2) the temperature of the steam from water boiling under a pressure of 76 cm. of mercury. These fixed points are called 0°C . and 100°C ., 32°F . and 212°F ., or 0°R . and 80°R .

Clinical, and maximum and minimum thermometers are special types of liquid in glass thermometers. Pyrometers, instruments for measuring high temperatures, are of two main types—thermo-electric and radiation pyrometers.

QUESTIONS

1. What is a scale of temperature? Explain carefully what is meant by a rise in temperature of 1°C .

Describe how you would make and graduate a mercury thermometer.

2. Explain:

(a) When a sensitive thermometer is placed in hot water the mercury first sinks, then rises.

(b) A clinical thermometer should never be cleansed in boiling water.

(c) The water in the radiators of motor-cars always boils on the Stelvio Pass (height 9040 ft.).

3. How does the sensitivity of a thermometer depend on (a) the volume of the bulb, (b) the cross-section of the stem, (c) the liquid?

4. Three boys read an ordinary thermometer, such as is used in the school laboratory, and record their readings as follows: 15°C ., 15.0°C ., 15.00°C . Which is correct and why?

5. If you were provided with a mercury thermometer, graduated in Centigrade degrees, and an ungraduated alcohol thermometer, how would you proceed to mark on the latter a scale of Fahrenheit degrees? (L.)

6. An ungraduated thermometer is attached to a cm. scale, and is found to read 7.3 cm. in melting ice, 23.8 cm. in steam at 100°C . and 3.5 cm. in a freezing mixture. What is the temperature of the freezing mixture?

7. Describe a form of thermometer that will record the minimum temperature during the night.

An ungraduated mercury thermometer attached to a millimetre scale reads 22.8 mm. in ice and 242.4 mm. in steam at standard pressure. What will it read on a day when the temperature is 22°F .? (C.)

8. What are the desirable properties of a perfect thermometric fluid? Show how far (a) mercury, (b) alcohol, (c) air, fulfil the conditions.

9. The freezing point of a Fahrenheit thermometer is correctly marked and the bore of the tube is uniform but it reads 76.5°F . when a standard Centigrade thermometer records 25° . What is the reading of the boiling point on this Fahrenheit thermometer? (L.)

10. Describe the clinical thermometer. How can such an instrument be graduated?

Some clinical thermometers are described as " $\frac{1}{2}$ minute thermometers", while others are called "1 minute thermometers". What do you think this means? How will two such thermometers differ from each other in construction?

11. A thermometer is required to measure temperatures between 10°C . and 20°C . to one-hundredth of a degree Centigrade. What features do you consider necessary? Will it be suitable for measuring rapidly varying temperatures between 10°C . and 20°C .? Give reasons for your answer.

12. What is the difference between "one degree Fahrenheit" and "one Fahrenheit degree"?

The temperature falls through 20°F . What is the corresponding fall in $^{\circ}\text{C}$.?

13. (i) What value on the Centigrade scale is equivalent to a temperature of (a) 185°F ., (b) 98.4°F ., (c) 0°F .?

(ii) What value on the Fahrenheit scale is equivalent to a temperature of (a) 15°C ., (b) 50°C ., (c) -273°C .?

(iii) What is the value of a temperature of 60°R . on the Centigrade and Fahrenheit scales?

14. On a certain day the readings on a Centigrade and Fahrenheit thermometer were the same. What was the temperature?

(C, 1st M.B.)

15. Describe the process of graduating a mercury thermometer whose range is from 0°F . to 250°F .

Some salt water is boiling in a flask. When the bulb of a Fahrenheit thermometer (which is graduated correctly) is placed in the water, the reading is 213°F . But when it is placed in the steam issuing from the water the reading is 211°F . Account for these facts. (L.)

16. How is it possible to measure extreme temperatures of a furnace or of liquid air? (N.)

Chapter II

THE EXPANSION OF SOLIDS AND LIQUIDS

Solids.

Heat causes bodies to expand. For instance, in the summer, telegraph wires are longer and therefore sag more than in the winter.

We can show that a metal bar expands when heated, by the following simple experiment. Clamp one end of a metal bar firmly and leave the other end free to roll over a small roller to which is attached a pointer (see Fig. 12). Heat the bar with a bunsen flame. It will expand and cause the pointer to move.

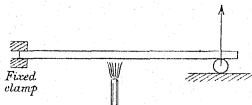


Fig. 12

We can show by similar experiments that the expansion of a bar

- (1) is proportional to the rise in temperature;
- (2) is proportional to the length of the bar;
- (3) varies with the nature of the substance of which the bar is composed.

Thus if we wish to compare the "expanding powers" of different substances we must find the expansion of equal lengths using an equal rise in temperature in each case. The **coefficient of linear expansion of a substance is the increase in length of unit length when its temperature is raised 1° .**

Now 1 cm. of iron when heated through 1°C. expands 0.000012 cm. Hence 1 inch, i.e. 2.54 cm., expands

$$2.54 \times 0.000012 \text{ cm.} = \frac{2.54 \times 0.000012}{2.54} \text{ inch} = 0.000012 \text{ inch.}$$

Similarly 1 mile expands 0.000012 mile. Any length expands by 0.000012 or $\frac{12}{1,000,000}$ ths of its original length.

Thus the coefficient of linear expansion represents the fractional increase in the length of a substance for a rise in temperature of 1° . The unit of length employed does not affect the value of the coefficient so long as the expansion is measured in the same units as the original length. A coefficient of linear expansion may therefore be regarded as a ratio and is usually expressed without units of length.

On the other hand, if an increase of temperature of 1° F. is taken instead of 1° C., the coefficient of expansion is only $\frac{5}{9}$ ths of its former value, since 1° F. is equivalent to $\frac{5}{9}^{\circ}$ C. We therefore write, 0.000012 per $^{\circ}$ C., or 0.0000067 per $^{\circ}$ F., showing which scale of temperature has been used.

Coefficients of Linear Expansion per $^{\circ}$ C.

Aluminium	0.000025	Glass (ordinary)	0.0000087
Brass	0.0000189	Pyrex glass	0.0000032
Copper	0.0000168	Invar	0.0000009
Iron and Steel	0.000012	Wood	0.000003-0.000005
Platinum	0.0000089		

The measurement of the coefficient of linear expansion.

The expansion of 1 cm. of a substance when heated through 1° C. is so small that in order to determine the coefficient of linear expansion we need to take a long bar and raise its temperature as much as is convenient. Even then the expansion is of the order of only a millimetre, and a metre scale is quite inadequate for measuring it accurately. Consequently we use a micrometer screw gauge.

There are many different forms of apparatus for finding the coefficient of linear expansion of a metal in the school laboratory. That shown in Fig. 13 was designed by Dr Barton of Repton School.

In principle it is very similar to the apparatus described at the beginning of this chapter. The expansion of a bar causes (by a simple mechanical arrangement) a pointer to move over a circular scale graduated in suitable divisions. The scale is calibrated by means of a micrometer screw gauge fixed at the other end of the bar.

First measure the length of the bar with a metre scale. Place the bar in position in the apparatus and turn the micrometer screw gauge, which has a pitch of $\frac{1}{2}$ mm., through exactly one turn. This causes the bar to move bodily through $\frac{1}{2}$ mm. Note the movement of the pointer, and hence obtain the number of divisions on the circular scale equivalent to an expansion of $\frac{1}{2}$ mm.

Turn the micrometer screw back through one turn, note the reading of the pointer and the temperature of the bar.

The bar is actually a hollow tube and it is heated by passing steam through it. When it has ceased to expand read the position of the pointer and the temperature of the bar.

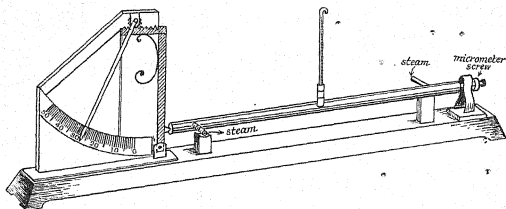


Fig. 13

The following readings were obtained using a copper bar:

Length of bar = 49.8 cm.

Calibration of pointer-scale.

No. of divisions on pointer scale $\equiv \frac{1}{2}$ mm. = 25.0.

Expansion of bar.

1st temp. of bar = 17.8° C.

2nd " " " = 98.3° C.

\therefore Rise in temp. of bar = 80.5° C.

1st reading of pointer = 2.9 divisions

2nd " " " = 36.2 "

\therefore Expansion of bar = 33.3 "

$$= \frac{33.3}{25.0} \times \frac{1}{2} \text{ mm.}$$

$$= 0.666 \text{ mm.}$$

49.8 cm. heated through 80.5°C . expand 0.0666 cm.

$$\begin{aligned} 1 \text{ cm. heated through } 1^{\circ}\text{C. expands } & \frac{0.0666}{49.8 \times 80.5} \\ & = 0.0000166 \text{ cm.} \end{aligned}$$

\therefore Coefficient of linear expansion of copper = 0.0000166 per $^{\circ}\text{C}$.

The bar-breaker.

Enormous forces are called into play when a bar expands or contracts. The bent and twisted girders in a large building which has been gutted by fire are convincing evidence of this.

An apparatus, known as the bar-breaker, is often used as a demonstration. A thick iron bar, *B*, rests horizontally in a massive iron stand as shown in Fig. 14. At one end of the bar is

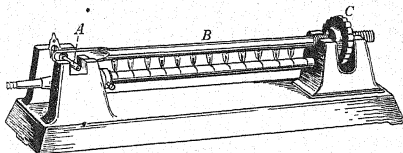


Fig. 14

a hole through which passes a small cast-iron bar, *A*, resting in a slot. The other end, *B*, is threaded and carries a stout nut, *C*, which can be screwed up. *B* is heated by means of a row of burners and, when it is hot, *C* is screwed up. As *B* contracts it exerts a force of the order of several tons and snaps the iron bar *A*.

COMPENSATIONS FOR AND APPLICATIONS OF EXPANSION

Surveying.

The first accurate determinations of coefficients of linear expansion were made in connection with surveying. In 1784 General Roy began the first ordnance survey of England. Only one distance required to be measured with a measuring rod, a base line of some 5 or 6 miles on Hounslow Heath. Starting

from this base line the whole country was divided into triangles with prominent landmarks as vertices. The angles of the triangles were measured by means of a theodolite and all the sides could be calculated using the length of the base line. It was of great importance, therefore, that the length of the base line should be known accurately. As its measurement took several months, General Roy, with the aid of Prof. Ramsden, performed a number of experiments to enable him to calculate the variations in the length of his measuring rods with change of temperature.

Recently M. Guillaume, of Paris, has invented a steel alloy containing 36 per cent. of nickel which has a negligible expansion. He has called it *invar* (the invariable) and measuring rods are now made of this alloy.

Steel gauges used in machine shops which are required to measure to one-thousandth of an inch are often marked "Correct at 60° F." and both they and the part to be measured must be at this temperature.

Expansion in structures.

Railway lines suffer considerable changes of temperature, being subject to direct sunshine in the summer and frost and

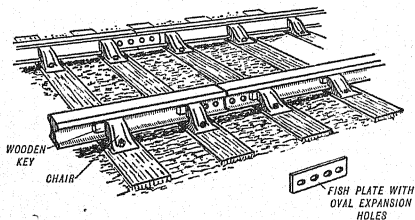


Fig. 15. Expansion joint in railway line.

snow in the winter. They are therefore laid down in strips with gaps between to allow for expansion. The strips are connected by fish plates with elongated holes to permit longitudinal movement (see Fig. 15).

A large steel-framed roof, such as that of a railway station, contracts and expands with changes of temperature, and would draw the walls in and out if it were attached rigidly to both of them. Consequently only one side of the roof is fixed and the other side rests either on rocker or roller expansion bearings (see Fig. 16), with sufficient space for movement.

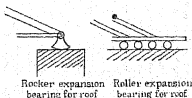


Fig. 16

In bridges, too, careful allowance must be made for expansion. The Forth bridge, which, with its approaches, is $1\frac{1}{2}$ miles long, has a total allowance of 6 ft. for



Fig. 17

expansion. Fig. 17 shows where this allowance is made. At *A* and *G* the ends of the outer cantilevers rest on roller bearings similar to those just described for a roof. At *C* and *E*, where the two central girders join the middle cantilever, are rocking posts.

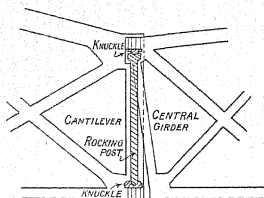


Fig. 18

Fig. 18 shows a rocking post. The central girder rests on the cantilever through a rocking post which can "rock", that is, tilt

to either side, when the central girder and cantilever expand or contract.

It is worth noting that there are other complicated arrangements in the design of the bridge to allow for a difference in temperature between its two sides.

Clocks and watches.

A clock is controlled by a pendulum whose time of swing depends on its length. If in the summer the pendulum expands, the clock will go more slowly and "lose"; similarly, in the winter, the clock will "gain".

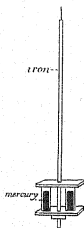


Fig. 19



Fig. 20

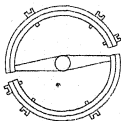


Fig. 21

In 1716 Graham invented a compensated pendulum known as the mercurial pendulum (see Fig. 19). The bob is hollow, and contains enough mercury to compensate by its upward expansion for the downward expansion of the iron rod. Thus when the temperature of the pendulum varies the position of its centre of gravity remains unchanged.

A modern pendulum compensated on the same principle consists of a bar of invar whose expansion is very small and hence may be compensated by a small bob of some common metal (see Fig. 20).

Watches and chronometers are controlled by a balance wheel and a steel hair-spring. Steel becomes less elastic as its temperature rises and the weakening of the spring causes the watch to

slow down. The expansion of the balance wheel has a further but much smaller tendency to make the watch go more slowly.

Compensation is effected by making the rim of the balance wheel of two metals, brass and steel (see Fig. 21). The more expansible metal, brass, is placed on the outer side of the rim. When the temperature rises the brass expands more than the steel, and causes the rim to curve inwards. The rim is usually weighted and the movement of the weight towards the centre causes the rate of swing to increase, thereby compensating for the loss of elasticity of the hair-spring. Open your watch, and see if it is compensated in this way.

A more modern and cheaper method of compensation is to make the hair-spring of *elinvar* (another of Guillaume's nickel-steel alloys) which retains its elasticity as the temperature rises.

By his invention of the composite bar Harrison was able to make an accurate chronometer and win a prize of £20,000 offered by Parliament, who were concerned to find an accurate method of determining longitude at sea. The mariner sets his chronometer by Greenwich time and observes its time at noon, i.e. when the sun is directly overhead. If his noon is 3 hours behind that of Greenwich, he knows that he is $\frac{1}{8}$ th of the way round the world. Suppose in the time taken to travel $\frac{1}{8}$ th of the way round the world, i.e. about 3000 miles, the chronometer gains or loses 1 minute, the error in position which results is $\frac{3000}{3 \times 60} = 17$ miles. In the days of sailing ships an uncompensated chronometer might well be an hour out at the end of so long a voyage. The error in this case was 1000 miles!

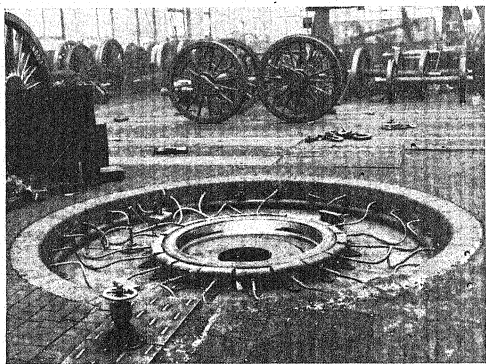
Applications of expansion.

Advantage is taken of expansion in riveting together the plates of a boiler. The rivets are put in red hot and when they contract pull the plates very tightly together.

The rim of a cart, train or locomotive wheel is always made so that it is slightly too small when cold. It is then heated and forced on while hot. In this way it is made to bind the wheel tightly. For a similar reason, the tubes of which the barrels of big guns are made are fitted over each other while hot.

Shrinking on, as this process is called, may also be done by cooling one of the parts in solid carbon dioxide ("dry ice"). The

shaft of a wheel, for example, may be cooled and thrust into the hole in the wheel: on warming up it expands and so grips. The method is employed in a car factory in the U.S.A. for setting special ring seatings for valves.



By courtesy of the Great Western Railway

Fig. 22. A gas tyre-heating furnace. The steel tyre of the driving wheel of a locomotive is forced on hot: when it cools it shrinks and binds the wheel tightly.

The expansion of glass.

It is well known that if boiling water is poured into a thick glass tumbler, the glass cracks. The inside of the tumbler expands, but since glass is a bad conductor of heat the outside remains cold and does not expand. The consequent strain causes cracking.

Silica and pyrex glass have a very low coefficient of expansion, and in their case very little strain is set up. They can be heated to red heat, and plunged into cold water without cracking.

Platinum alone of the metals has a coefficient of linear expansion equal to that of glass. It has therefore been much used for

sealing into glass. One of M. Guillaume's nickel-steel alloys, however, with the requisite coefficient of expansion, is now taking its place for this purpose, and is known as platinum substitute.

Superficial expansion.

• Superficial expansion is an expansion of area. When a bar or plate is heated its breadth increases as well as its length.

The coefficient of superficial expansion is the increase in area of unit area when heated through 1° .

Many substances, known as isotropic substances, have an equal coefficient of linear expansion in the direction of their breadth (and also their depth), as of their length. There is no need to make a separate determination of the coefficient of superficial expansion for isotropic substances, since it is double the linear coefficient. (Glass and the metals are examples of isotropic substances, while wood and many crystals are non-isotropic.)

Consider a plate of metal 1 cm. square. Suppose it is heated through 1° C.

Let α = coefficient of linear expansion (per $^{\circ}$ C.).

• The length and breadth both increase from 1 cm. to $(1 + \alpha)$ cm.

• \therefore New area = $(1 + \alpha)^2$.

• \therefore Increase in area = $(1 + \alpha)^2 - 1$
 $= 1 + 2\alpha + \alpha^2 - 1$
 $= 2\alpha + \alpha^2$.

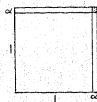


Fig. 23

(See shaded area in Fig. 23.)

Suppose the plate in question is made of iron, here

$$2\alpha = 0.000024,$$

$$\alpha^2 = 0.00000000144.$$

It is clear that we can ignore α^2 compared with 2α since in comparison it is negligibly small. Thus the increase in area of 1 sq. cm. when heated through 1° C. is 2α ; in other words, the coefficient of superficial expansion is twice the coefficient of linear expansion.

Cubical expansion.

Substances when heated expand in all three directions, length, breadth and height. They experience a change of volume, known as cubical expansion.

The coefficient of cubical expansion is the increase in volume of unit volume when heated through 1° .

Prove, by the method of the preceding article, that for isotropic substances the coefficient of cubical expansion is 3α , i.e. three times the coefficient of linear expansion.

The expansion of liquids.

Liquids have no shape and therefore no fixed dimensions. Consequently, in their case, we are concerned only with cubical expansion.

Moreover, they have always to be held in some containing vessel. Owing to the expansion of the vessel the apparent expansion of the liquid is less than the true value.

Two coefficients of expansion are used.

The coefficient of apparent expansion is the apparent increase of unit volume per 1° rise in temperature (no allowance being made for the expansion of the containing vessel).

The coefficient of absolute expansion is the true increase in unit volume per 1° rise in temperature (allowance being made, if necessary, for the expansion of the containing vessel).

There is a simple relation between these two coefficients.

Coefficient of absolute expansion of liquid = coefficient of apparent expansion of liquid + coefficient of cubical expansion of material of containing vessel.

DETERMINATION OF COEFFICIENTS OF EXPANSION OF LIQUIDS

1. *Volume dilatometer.*

To determine the coefficient of apparent expansion of a liquid we may use a volume dilatometer, i.e. a flask of known volume with a graduated stem (see Fig. 24).

Pour the liquid into the dilatometer so that its surface level is at the lower end of the stem. Read its temperature and volume. Heat the dilatometer for some considerable time in a water-bath and again read the temperature and volume of the liquid.

Hence calculate the (apparent) increase in volume of 1 c.c. when heated through 1° C.

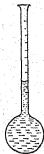


Fig. 24

2. The weight thermometer method.

The most accurate method of determining volumes is by weighing, since the balance is the most accurate instrument in the laboratory.

• Weigh a density bottle, fill it with the liquid to the top of the stopper and reweigh. Take the temperature of the liquid. Heat the bottle and its contents in a water-bath, for about ten minutes, until the liquid has attained the temperature of the water-bath. Read the temperature. Some of the liquid will have been expelled from the bottle: mop any drops from the top of the stopper with blotting paper to prevent them from being drawn in when the liquid cools. Cool and dry the bottle and liquid, and reweigh.

The following readings were obtained using glycerine:

Weight of empty bottle	= 17.33 gm.
„ „ bottle and liquid at 15.6° C.	= 81.84 „
„ „ „ „ „ 100.0° C.	= 79.13 „
∴ „ „ liquid filling bottle at 15.6° C.	= 64.51 „
• „ „ „ „ „ 100.0° C.	= 61.80 „

To calculate the coefficient of apparent expansion of the liquid we require to know its original volume and its expansion. We take the weight of the liquid left in the bottle at the end of the experiment (see Fig. 25) to represent the original volume, and the weight of the liquid in the bottle at the beginning of the experiment to represent its final volume. The difference in these weights represents the expansion.



Fig. 25

This is perfectly logical, since if we start again with the liquid left in the bottle at 15.6° C. and reheat it to the final temperature, it will expand to fill the bottle exactly. The real point is that in comparing volumes by means of the weights of a liquid, the liquid must be at the same temperature in each case (here at 15.6° C.). For volume changes with temperature, but weight does not.

Thus 61.80 volumes heated through 84.4° C. expand 2.71 volumes.

\therefore 1 volume heated through 1°C . expands $\frac{2.71}{61.8 \times 84.4}$ volumes
 $= 0.000520$ volumes.

\therefore Coefficient of apparent expansion of liquid $= 0.00052$ per $^{\circ}\text{C}$.

A correction.

In the determination of the coefficient of cubical expansion of a liquid the original volume of the liquid depends on the initial temperature. Hence if the experiment is repeated starting at a different temperature, a slightly different result will be obtained. In accurate work it is usual to convert the original volume of the liquid to its volume at 0°C . Using the readings in the weight thermometer experiment, the correction is made as follows.

Expansion in 84.4°C . $= 2.71$ volumes.

\therefore Contraction in 15.6°C . $= \frac{2.71 \times 15.6}{84.4}$ volumes

$= 0.50$ volumes.

\therefore Volume at 0°C . $= 61.80 - 0.50$ volumes

$= 61.30$ volumes.

\therefore Coefficient of cubical expansion $= \frac{2.71}{61.3 \times 84.4}$
 $= 0.000524$ per $^{\circ}\text{C}$.

Except in very accurate work this correction is less than the margin of experimental error.

We may here mention a well-known formula which is derived from the definition of the coefficient of cubical expansion.

Let V_0 = volume of substance at 0°C .

V_t = " " " $t^{\circ}\text{C}$.

γ = coefficient of cubical expansion.

By definition γ = increase in volume of unit volume (at 0°C .) per 1°C . rise

$$= \frac{V_t - V_0}{V_0 t}.$$

$$\therefore V_t = V_0 (1 + \gamma t).$$

The effect of expansion on density.

When a substance is heated its volume increases but its weight is unchanged.

$$\text{Now Density} = \frac{\text{Weight}}{\text{Volume}}.$$

Hence, on heating, the density of a substance decreases.

Let V_0 = volume of substance at 0°C. ,

V_t = " " " " $t^\circ \text{C.}$,

W = weight of substance,

γ = coefficient of cubical expansion.

$$\text{We know} \quad V_t = V_0 (1 + \gamma t) \quad (\text{see p. 28}).$$

$$\therefore \frac{V_t}{W} = \frac{V_0}{W} (1 + \gamma t).$$

$$\text{But} \quad \frac{W}{V_0} = \text{density at } 0^\circ \text{C.} = \rho_0,$$

$$\frac{W}{V_t} = \text{density at } t^\circ \text{C.} = \rho_t,$$

$$\therefore \frac{1}{\rho_t} = \frac{1}{\rho_0} (1 + \gamma t)$$

$$\text{or} \quad \rho_0 = \rho_t (1 + \gamma t).$$

Direct determination of the absolute coefficient.

Advantage may be taken of the change in density of a liquid due to expansion, to find its absolute coefficient of expansion direct.

There is a well-known method of comparing the densities of two liquids by balancing them in the arms of a U-tube. Dulong and Petit, in 1817, balanced a hot column of mercury against a cold column. One arm of a U-tube containing mercury is surrounded by a cold water jacket at $t^\circ \text{C.}$ and the other arm by a steam jacket at 100°C. (see Fig. 26). When conditions are steady the heights of the two columns h_t, h_{100} are measured.

Now, according to the simple laws of liquid pressure,

$$\frac{\text{Density of liquid in left column}}{\text{Density of liquid in right column}} = \frac{\text{Height of right column}}{\text{Height of left column}}.$$

$$\therefore \frac{\rho_t}{\rho_{100}} = \frac{h_{100}}{h_t}.$$

But we have shown in the last article that

$$\frac{1}{\rho_t} = \frac{1}{\rho_0} (1 + \gamma t)$$

and similarly
$$\frac{1}{\rho_{100}} = \frac{1}{\rho_0} (1 + \gamma \times 100).$$

Dividing
$$\frac{\rho_t}{\rho_{100}} = \frac{1 + 100\gamma}{1 + \gamma t}.$$

$$\therefore \frac{1 + 100\gamma}{1 + \gamma t} = \frac{h_{100}}{h_t}.$$

Hence knowing h_{100} , h_t , and t , γ can be calculated.

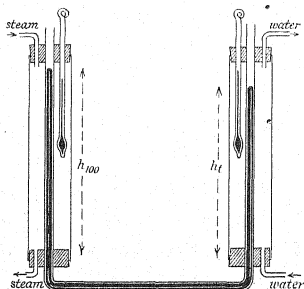


Fig. 26

The cross-sections of the two arms of the U-tube need not necessarily be equal since the pressure of a liquid depends only on its depth and not its cross-section. Thus the expansion of the arm surrounded by the steam-jacket has no influence on the result: it is the coefficient of *absolute* expansion that is obtained.

The determination of the absolute coefficient of expansion of a liquid, such as mercury, in this way, enables the coefficient of cubical expansion of the glass of a weight thermometer to be

found. How? The result can be used in the determination of the coefficient of absolute expansion of other liquids with that weight thermometer.

The peculiar expansion of water.

When water is cooled it contracts like normal substances until it reaches 4°C . Then on cooling from 4°C . to 0°C . it *expands*. The graph in Fig. 27 enables one to see at a glance this anomalous behaviour. What would the corresponding graph for mercury be like?

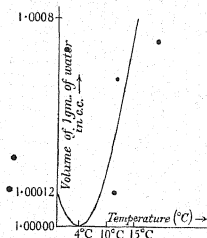


Fig. 27

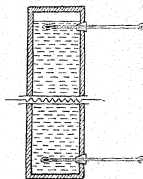


Fig. 28

The anomalous expansion of water can be demonstrated by a modification of Hope's apparatus due to Mr D. G. A. Dyson.

Ice-cold water contained in a cylindrical glass vessel (well insulated by cork to minimise gain of heat from the surroundings) is heated at its middle (see Fig. 28) by an electric heater. The temperatures near the top and bottom of the water are taken at intervals by the thermometers shown.

The water in the middle, on being heated, at first *contracts* and hence sinks: when its temperature has reached 4°C ., on further heating it begins to expand. Soon its density is less than that of the water above it at 0°C . and hence it rises. The reading of the upper thermometer now rises steadily while that of the lower remains steady at 4°C . (see Fig. 29).

Hope's method was to cool water originally at room temperature by a freezing mixture placed round the outside of the

middle of the vessel. Draw the graphs to illustrate the reading of the two thermometers in this case. The advantage of the new method is that it takes about ten minutes instead of over an hour.

The anomalous expansion of water seems almost a beneficent dispensation of Providence. For without its operation, the aquatic life in many of our ponds and lakes would perish in a severe winter. Since water has a maximum density at 4°C ., the

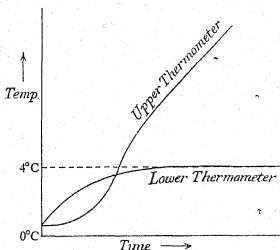


Fig. 29

temperature at the bottom of a lake which has frozen solid on the surface remains at this value. Cooling by circulation of the water, known as convection, ceases at 4°C ., since the heaviest liquid now remains at the bottom.

SUMMARY

The coefficient of $\left\{ \begin{array}{l} \text{linear} \\ \text{superficial expansion of a substance} \\ \text{cubical} \end{array} \right.$
 is the increase in $\left\{ \begin{array}{l} \text{length} \\ \text{area} \\ \text{volume} \end{array} \right.$ of unit $\left\{ \begin{array}{l} \text{length} \\ \text{area} \\ \text{volume} \end{array} \right.$ when its temperature is raised 1° .

In the case of isotropic substances, which have the same linear coefficient of expansion, α , in all directions,

Coefficient of superficial expansion = 2α ,

„ „ cubical „ = 3α .

* Compensations for expansion are made in metal structures, clocks and watches.

Two coefficients of cubical expansion are used in the case of liquids:

1. The coefficient of **apparent expansion**, in which no allowance is made for the expansion of the containing vessel.

2. The coefficient of **absolute expansion**, in which allowance is made, if necessary, for the expansion of the containing vessel.

Coefficient of absolute expansion of liquid

= Coefficient of apparent expansion of liquid

+ Coefficient of cubical expansion of material of containing vessel.

The coefficient of apparent expansion of a liquid may be determined by means of a dilatometer or weight thermometer.

$$V_t = V_0 (1 + \gamma t)$$

and

$$\rho_0 = \rho_t (1 + \gamma t) \quad (\text{see p. 29}).$$

* The coefficient of absolute expansion of a liquid may be determined direct by the U-tube method of Dulong and Petit.

When the temperature of water is raised from 0°C. to 4°C. it contracts; from 4°C. upwards it expands. Hence water has a maximum density at 4°C.

QUESTIONS

1. It is stated that the coefficient of linear expansion of brass is 0.0000189 per $^\circ \text{C.}$ What is meant by this statement? Describe an accurate method of determining the coefficient of linear expansion of brass. (N.)

2. Find the expansion of a brass rod 100 cm. long when heated from 15°C. to 100°C. (see table, p. 17).

3. By how many inches will a steel bridge of 800 ft. span increase in length when its temperature rises from 0°F. to 100°F. ? (Coefficient of linear expansion of steel = 0.000011 per $^\circ \text{C.}$) (O. & C.)

4. A railway line is laid with 40 ft. lengths of rail on a day when the temperature is 10°C . and gaps of $\frac{1}{4}$ in. are left between successive lengths. At what temperature would the gaps close up? (Coefficient of linear expansion of steel = 0.000012 per $^{\circ}\text{C}$.) (C.)

5. If the graduation on a brass scale is correct at 0°C . what would be the true length of an object that was read by the scale at 25°C . as 30.15 in.? (Coefficient of linear expansion of brass = 0.000018 per $^{\circ}\text{C}$.) (L.)

6. Give two examples of devices used to overcome difficulties produced by the expansion of substances by heat. Explain the principles on which they are based.

Iron rods AB and AC , 3 and 4 cm. long, respectively, are jointed at right angles at A . What must be the coefficient of expansion of the material forming the side BC of the triangle if it is to remain right angled at higher temperatures? (Coefficient of linear expansion of iron = 0.000011 .) (O.)

7. A rod 1000 cm. long at 0°C . is composed of two metal rods A and B joined end to end, the two portions having coefficients of linear expansion 0.00005 and 0.00002 respectively. Find the lengths of the two portions A and B of the rod if the expansion is 3 cm. at 80°C . (O.)

8. Find the ratio of the lengths of two rods of different metals A and B at 0°C . if the difference of their lengths is the same at all temperatures. The coefficients of linear expansion of A and B may be taken as 0.00002 per $^{\circ}\text{C}$. and 0.00004 per $^{\circ}\text{C}$. respectively. (C.)

9. A man wishes to fit an aluminium ring on a rod of iron of 1 in. diameter, but it is 0.001 in. too small in diameter. How much should its temperature be raised before it will just slip on? Subsequently he wishes to remove it again, but now he has to heat both metals together. Through how many degrees must this be done? (Coefficient of expansion of aluminium 0.000025 ; coefficient of expansion of steel 0.000010 .) (N.)

10. Suggest reasons why gaps, similar to those in railway lines, are not left in tramlines.

11. Explain:

(a) If a hot lamp glass be touched with a cold knife it will probably crack.

(b) A steel or copper wire cannot be sealed through glass.

(c) A stopper can sometimes be loosened by heating gently the neck of the bottle.

(d) A dent can often be removed from a ping-pong ball by holding it in front of a fire.

(e) A cheap watch will not keep good time throughout the year.

12. Why must clocks and watches be compensated for changes in temperature? Describe, with diagrams, some form of compensated (a) pendulum, (b) balance wheel.

13. Fig. 30 shows an old form of compensated pendulum known as Harrison's gridiron pendulum, consisting of alternate rods of steel, *S*, and brass, *B*. The steel rods expand downwards and the brass upwards. Using the table on p. 17 find the ratio of the total lengths of the steel and brass rods. Why, do you think, are more than two rods used?

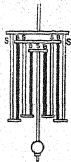


Fig. 30

14. Show how the coefficient of cubical expansion of a homogeneous substance is related to its coefficient of linear expansion.

Describe a laboratory method of determining the coefficient of apparent expansion of mercury in glass. (C.)

15. The height of the mercury in a barometer is 755 mm. at 10°C . What would be its height if the temperature were 25°C ? (Coefficient of expansion of mercury = 0.00018 per $^{\circ}\text{C}$.)

16. A litre bottle full of milk which has just been pasteurised at 60°C . is sealed and cooled to 15°C . Find the volume of the empty space in the bottle at the new temperature, assuming that the coefficients of cubical expansion of milk and glass are 0.00038 and 0.000025 per $^{\circ}\text{C}$. respectively. (O. & C.)

17. When calculating the coefficient of apparent expansion of a liquid from the readings of an experiment with a density bottle the expansion of the liquid left in the bottle must be considered. Explain carefully why this is so. What would be the actual source of error if the weight of the liquid originally filling the bottle were taken to represent the volume of the liquid before its expansion?

18. Define the coefficients of apparent and real expansion of a liquid. What is the relation between them?

A glass vessel contains when full 816.00 gm. of mercury at 0°C . The mass of mercury which fills it at 100°C . is 803.21 gm. The coefficient of absolute expansion of mercury is 0.000182 per $^{\circ}\text{C}$. Calculate the coefficient of cubical expansion of glass. (C.)

19. A glass vessel which holds exactly 1000 gm. of mercury at 15°C . is put into boiling water at 100°C . What weight of mercury is expelled? (Coefficient of dilation of mercury = 0.00018 ; coefficient of expansion (linear of glass) = 0.00001 .) (N.)

20. A thermometer has a bulb of volume 0.5 c.c. and the mercury in the stem moves 2 mm. for 1°C . rise in temperature. If the coefficient of cubical expansion of mercury is 0.00018 per $^{\circ}\text{C}$., calculate approximately the diameter of the circular bore of the stem, neglecting the effect of the expansion of the glass. (N.)

21. A mercury-in-glass thermometer is graduated in degrees Centigrade. The internal area of cross-section of the stem is 0.05 sq. mm. and the internal volume of the bulb and stem up to the 0° mark is 0.8 c.c. Assuming the coefficient of cubical expansion of mercury to be 0.00018 and neglecting the expansion of the glass, calculate the distance between the 0° and 100° marks. (O. & C.)

22. If the two arms of a U-tube are kept at different temperatures the levels of any contained liquid are not the same in the two arms. Explain why this is so and describe how it may be made the basis of a method of measuring the coefficient of expansion of the liquid. Will the coefficient of expansion so obtained be real or apparent? (O. & C.)

23. The density of mercury at 0°C . is 13.60 gm. per c.c. and its coefficient of expansion is 0.00018 per $^{\circ}\text{C}$. Find its density at 100°C .

24. Show how the density of a substance at different temperatures is connected with its coefficient of expansion.

A U-tube containing a liquid has the two limbs maintained at 15°C . and 100°C . respectively. On reaching a steady state the lengths of the liquid columns are 97 cm. and 102 cm. What is the coefficient of expansion of the liquid? Is this the real or the apparent coefficient? Give reasons. (L.)

25. A copper cylinder weighs less in water than in air. How would this difference in weight be affected by raising the temperature of the water? Give reasons for your answer. (L.)

26. Describe an experiment to show that water has a maximum density at 4°C . What important consequences follow from this peculiar property of water?

Will a water (in glass) thermometer show a minimum reading at 4°C . or at a lower or higher temperature than this? Give reasons for your answer. (L.)

27. Under what circumstances is the coefficient of expansion of water (a) negative, (b) zero?

28. Some ice at -10°C . is heated to 50°C . Plot a graph to show the change in volume of the system over the range of temperature and discuss the natural consequences which follow from the shape of the graph.
(O. & C.)

29. A tall cylindrical jar of water at 10°C . is surrounded about its middle by a freezing mixture, and thermometers are inserted through the sides near top and bottom. Explain how cooling proceeds and what happens to the thermometer readings and why? What difference would there be if, instead, the jar were cooled by dropping ice in gently (the freezing mixture being removed)?

Chapter III

THE EXPANSION OF GASES

The volume of a gas may be increased not only by raising its temperature, but also by reducing its pressure. There are thus three dependent variables, temperature, volume, and pressure. It will be remembered that in the case of the expansion of solids and liquids there are only two dependent variables, temperature and volume.

The simplest method of investigating the expansion of gases is to keep one of these variables constant, and find the relation between the other two. The three variables are kept constant each in turn and the relations found between

1. pressure and volume when the temperature is kept constant;
2. volume and temperature when the pressure is kept constant;
3. pressure and temperature when the volume is kept constant.

The relation between the volume and pressure of a gas at constant temperature.

Robert Boyle. The first investigations into the compressibility of gases were made by Robert Boyle. Boyle was born in Ulster in 1627, the seventh son of the Earl of Cork. He was educated at Eton and also under a tutor during a term abroad. While at Geneva he was deeply impressed by a severe thunderstorm; "every clap was both preceded and attended with flashes of lightning, so frequent and so dazzling that he began to imagine them the sallies of that fire that must consume the world". This experience turned his thoughts to religion for the rest of his life. He was a great experimenter and wrote: "In my laboratory I find that water of Lethe which causes that I forget everything but the joy of making experiments."

Boyle's Law.

Boyle's first scientific investigations, performed in his private laboratory at Oxford, were concerned with the pressure of the atmosphere. News of the recent experiments of Torricelli,

Pascal, and von Guericke had reached him and he decided to repeat them for himself.

It was probably a fantastic claim made by Franciscus Linus, a Jesuit of Liège, that caused him to devise his experiments on the compressibility of air. It is now a known fact that a column of mercury about 29 in. high can be supported in a barometer tube by the pressure of the atmosphere. Linus scoffed at the idea of so great a weight of mercury being supported by air, and said that the actual support consisted of invisible threads running from the mercury to the top of the tube. These threads, he said, he had felt with his fingers!

Boyle was determined to prove that air, by its pressure or "spring", could support even greater columns of mercury than 29 in. He took a long glass tube sealed at one end and bent it so that the sealed arm was much shorter than the open arm (see Figs. 31*a* and *b*). He poured in mercury and measured the length of the trapped air column in the sealed arm when the mercury stood at the same height in each arm. The air was then at atmospheric pressure, equivalent to about 29 in. of mercury. He poured in more mercury until the difference in level in the two arms was 29 in. The air was now under a pressure of two atmospheres, and Boyle found, "not without delight and satisfaction", that its volume was exactly halved. On pouring in more mercury until its pressure was three atmospheres, he found that its volume had been reduced to one-third of its original value.

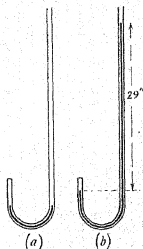


Fig. 31

The facts which he discovered enabled him to formulate the law that bears his name:

The pressure and volume of a fixed mass of gas are inversely proportional, provided that the temperature is kept constant.

I.e.

$$p \propto \frac{1}{v},$$

where p = pressure of gas, v = volume of gas, or

$$pv = \text{constant.}$$

Some years later Mariotte, a Frenchman, discovered the law independently, and it is known to every French schoolboy as Mariotte's Law.

A slight modification of Boyle's method is usually employed in the school laboratory when repeating his experiment. The two arms of the glass tube are connected at their lower ends by a rubber tube (see Fig. 32), and the pressure on the trapped air may be increased by raising the open arm or squeezing the rubber tubing.

The levels of *A*, *B* and *C* are read on a vertical scale. The volume of the gas is proportional to the length *AB* (assuming that the tube has a uniform cross-section). The pressure of the gas is equal to the atmospheric pressure (obtained by reading the barometer) *plus* the vertical height *CB* (cm. mercury). For the pressure of the atmosphere is equivalent to that of a column of mercury, about 76 cm. high, resting on *C*. Hence the pressure at the level *B* is equivalent to a column of mercury $76 + CB$ cm. long.

Thus if v = volume of gas in *AB*,

p = pressure of gas in *AB*,

$v \propto AB$,

$p \propto \text{barometer reading} + CB$ cm. mercury.

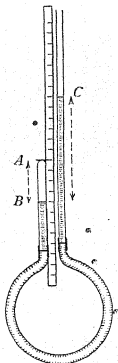


Fig. 32

Take several sets of readings and enter them in columns headed v and p . In a third column write down the product pv . What will you expect to notice about this quantity? (The method described in the author's *Hydrostatics and Mechanics* is another way of proving the same law.)

The Royal Society.

A fitting conclusion to our account of the researches of Robert Boyle is a mention of the founding of the Royal Society of which he was one of the first and most distinguished members. It was about the year 1645 that a number of scientific men began to meet together to discuss such subjects as "physick, anatomy, geometry, astronomy, navigation, staticks, (magnetics, chymicks,

mechanics, and natural experiments". They formed a kind of "invisible college" and their aim was the Advancement of Learning. They met sometimes in London and sometimes in Oxford, and among their earliest members were Sir Isaac Newton and Sir Christopher Wren. In 1660 they were granted a Royal Charter by Charles II and became fellows. Today Fellowship of the Royal Society is one of the greatest honours that can be conferred on a British scientist.

It is an interesting coincidence that about this time Mariotte became one of the first members of a kindred French society, the Académie des Sciences, which was founded in 1666.

The relation between the volume and temperature of a gas at constant pressure.

At the beginning of the nineteenth century accurate experiments on the expansion of gases were carried out independently by three Frenchmen—Charles, who became interested in the subject from his experiences as a pioneer balloonist, Gay-Lussac, and Regnault.

A convenient apparatus for investigating the relation between the volume and temperature of a gas at constant pressure is shown in Fig. 33. The gas is confined in the bulb *A* by means of concentrated sulphuric acid, which acts as a pressure indicator, and also as a drying agent. It is essential that the gas should be absolutely free from water vapour since the latter exerts an additional pressure which increases with temperature. Indeed the results of the earliest experiments performed in the eighteenth century differed widely owing to the non-realisation of this fact.

The volume of the gas is read by means of the graduations below *A* and its temperature is raised by the water-jacket which is electrically heated. Both glass and gases are bad conductors of heat and it is therefore necessary to keep the bath at one temperature for several minutes before taking a reading. Heat the bath a few

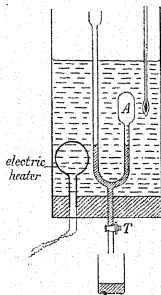


Fig. 33

degrees above the desired temperature and stir steadily. Level the acid in the two arms by running some out through the tap *T* in order to bring the pressure of the gas to atmospheric pressure. Read the temperature and volume of the air simultaneously. Take readings from room temperature about every 10°C. so that about seven readings are obtained. Record the results in two columns of a table. Plot volume against temperature.

The resulting graph will be a straight line, proving that the volume increases regularly with the temperature (as measured with a mercury thermometer). From the graph find the volume at 0°C. and also the increase in volume from say 0°C. to 100°C. , and hence the expansion per 1°C. Then:

$$\text{Coefficient of expansion} = \frac{\text{Increase in volume per } 1^{\circ}\text{C.}}{\text{Volume at } 0^{\circ}\text{C.}}$$

[This definition is exactly the same as the amended definition for a liquid—see p. 28. In the case of a gas it is essential to take the initial volume at 0°C.]

Express the result in the form of a fraction. The advantage of using the graph in this way is that it represents the average of all the readings.

The results of a series of experiments of this kind disclose the remarkable fact that *all* gases, unlike solids and liquids, have *the same* coefficient of expansion, namely $\frac{1}{273}$ per $^{\circ}\text{C.}$

The fact is usually embodied in what is known as Charles' Law.

All gases expand by $\frac{1}{273}$ of their volume at 0°C. for each 1°C. rise in temperature, provided that they are maintained at constant pressure.

The absolute scale of temperature.

The temperature -273°C. is believed by physicists to be the lowest which can ever be attained, and is known as the Absolute Zero.

One of the chief claims of evidence which led to this striking assumption was derived from the study of the expansion and contraction of gases.

Take a large sheet of graph paper and plot the readings of volume and temperature obtained in the experiment just described. Produce the graph backwards to cut the temperature

axis when the volume is 0 (see Fig. 34). This temperature will be found to be about -273°C . It is the temperature at which the gas would have contracted to zero volume if the latter were cooled and remained a gas; and it is the same for all gases. We shall see the true significance of this apparently impossible conclusion when we study the kinetic theory of gases (see p. 47).

Moreover as -273°C . is approached, further cooling becomes extraordinarily difficult, just as the last traces of air are most difficult to remove when producing a perfect vacuum. In spite of difficulties, however, and after twenty years of research, Prof. Onnes and his assistants at Leyden have succeeded in reaching a temperature within 0.1°C . of absolute zero.

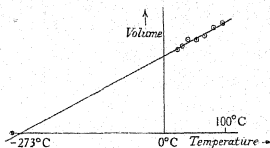


Fig. 34

We can now achieve what Fahrenheit attempted in vain—a scale of temperature with 0° as the lowest attainable temperature.

This scale, commencing at -273°C ., with Centigrade degrees, is known as the Absolute Scale.

$$\begin{aligned}\text{Thus} \quad -273^{\circ}\text{C.} &= 0^{\circ}\text{A.} \\ 0^{\circ}\text{C.} &= 273^{\circ}\text{A.} \\ 100^{\circ}\text{C.} &= 373^{\circ}\text{A.,}\end{aligned}$$

and so on.

The relation between the pressure and temperature of a gas at constant volume.

To investigate how the pressure of a gas depends on the temperature, at constant volume, the apparatus shown in Fig. 35 may be used. The gas under test is confined in the bulb *A* by mercury in the U-tube. It is kept at a constant volume by adjusting the mercury level to a fixed mark *B*. This is done by

raising or lowering the outer arm. The pressure of the gas in $A = \text{height of mercury barometer} + h$, where h is the difference in the levels of the mercury at C and B . The bulb A is heated in a water-bath.

Take readings of temperature and pressure about every 10°C. , from room temperature to 100°C. As the air in the bulb A takes some time to reach the temperature of the water-bath, heat the bath with a bunsen burner 2 or 3°C. above the required temperature and stir steadily for some minutes before bringing the mercury to the level B and reading the level C simultaneously with the temperature.

Read the barometer.

Plot the temperature against pressure.

If your graph is a straight line what does this indicate? From the graph find the pressure at 0°C. and also the increase in pressure per 1°C. rise. Hence find

$$\frac{\text{Increase in pressure per } 1^\circ \text{C. rise}}{\text{Pressure at } 0^\circ \text{C.}},$$

which is called the *pressure coefficient* (compare it with coefficient of expansion).

Draw a graph on a large sheet of graph paper and find at what temperature the gas should exert no pressure.

As a result of a series of experiments of this kind, it can be shown that, when gases are heated at constant volume, the increase in their pressure obeys a law very similar to that of Charles.

The pressure of all gases increases by $\frac{1}{273}$ of their pressure at 0°C. for each 1°C. rise in temperature, provided that their volume is kept constant.

The constant-volume gas thermometer.

The ordinary mercury thermometer is a very imperfect instrument owing to the vagaries of its glass envelope. It has been found, for instance, that the glass bulb of a thermometer shrinks slowly with age, causing a raising of the lower fixed point. Joule determined the "zero reading" of one of his thermometers

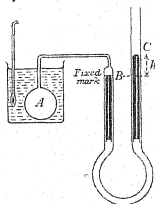


Fig. 35

at intervals over a period of 38 years. The following is a list of some of his readings:

Apr. 1844	0° F.	Apr. 1856	0.73° F.
Feb. 1846	0.42° F.	Feb. 1873	0.91° F.
Jan. 1848	0.51° F.	Dec. 1882	1.02° F.

It was realised that owing to the unreliability of mercury thermometers there were serious sources of error in the early researches in heat. It was clear therefore that some form of standard thermometer, against which mercury thermometers could be calibrated at intervals, was an urgent necessity.

The standard decided upon was a gas thermometer. Since gases have a much larger coefficient of expansion than mercury, slight irregularities in the volume of the bulb are of less importance. Moreover, since a gas such as hydrogen does not liquefy until -253°C. , gas thermometers have a very much wider range than liquid-in-glass thermometers.

It is clear that there are two possible types of air thermometer, utilising either the increase in volume of a gas at constant pressure, or the increase in pressure of a gas at constant volume, to measure temperature. The latter was chosen in preference to the former since it is easier to keep the volume of a gas constant than its pressure, owing to variations in atmospheric pressure.

Thus the apparatus shown in Fig. 35 represents a simple constant volume gas thermometer. The instrument must first be calibrated by finding the pressures at the fixed points. Then to obtain a temperature, say of water, the bulb is immersed in the water and the pressure of the gas read. The temperature corresponding to this pressure may then be calculated.

Conversion of gas volumes.

A common need of the chemist is to calculate what volume a certain mass of gas would have if its temperature and pressure were changed. As a rule he wishes to know the volume at normal temperature and pressure (N.T.P.), i.e. 0°C. and 76 cm. of mercury. Here we have an application of the laws of Boyle and Charles.

Example. To convert a volume of 50 c.c. of gas from a pressure of 75 cm. to 76 cm. of mercury.

At a pressure of 75 cm. the volume is 50 c.c.

The two chief characteristics of a gas are its compressibility and its springiness. A gas always exerts a pressure and tends to expand.

We can account for these characteristics by assuming that a gas consists of tiny particles, called molecules, incessantly moving, and with large empty spaces between them. When a gas is compressed the molecules are pushed more closely together and the empty space is reduced. The tendency of a gas to expand is due to the motion of the molecules, and the pressure it exerts to the bombardment of the molecules on the walls of its containing vessel.

The theory is known as the Kinetic Theory of Gases.

The Gas Laws and the Absolute Zero.

Consider a gas in a cylinder, closed by a piston. The molecules of the gas will bounce on the piston and exert a pressure on it. If the piston is moved in so that the volume of the gas is halved, twice as many molecules will hit the piston per second and hence the pressure on it will be doubled. In this way we can explain Boyle's Law.

Suppose the gas in the cylinder is heated and the volume is kept constant by keeping the piston fixed. The pressure increases. How are we to account for this increase in pressure? The velocity of the molecules must have increased. Thus *raising the temperature of a gas means increasing the velocity of the molecules.*

If a constant external pressure is applied to the piston and the gas is heated, the piston will be forced out until the reduction in the number of impacts of the molecules on the piston counteracts their increase in speed. That is, if a gas is heated at constant pressure it expands.

When a gas is cooled the velocity of its molecules decreases. We can conceive that ultimately the molecules will stop moving altogether. This will occur at the temperature of Absolute Zero, -273°C . At this temperature the gas will exert no pressure and its volume will be very small—that of the molecules only.

Some statistics.

The Kinetic Theory of Gases, drawing evidence from many other sources besides the gas laws, has been very fully developed quantitatively and mathematically. We can only present here,

however, without going into their derivation, the following extraordinary statistics:

In 1 c.c. of air at N.T.P. there are 27×10^{18} molecules, i.e. twenty-seven million billion. They travel at about 1000 m.p.h. and as there are so many of them moving at such a high speed they make an enormous number of collisions, about 10^{10} , i.e. ten thousand million per second. Since gases do not settle like a cloud of dust in a layer upon the floor, we know that the molecules must be perfectly elastic. A perfectly elastic golf ball would always bounce to the same height from which it was dropped, and hence, if left, go on bouncing (*in vacuo*) for ever. This is an impossibility in the case of a golf ball but it must be true in the case of the molecules.

Since gases are so compressible the molecules must be very tiny compared with their distance apart. Air for instance can be compressed to something like $\frac{1}{200}$ th of its normal volume, proving that the empty space in air at atmospheric pressure must be at least 200 times the space occupied by the molecules.

We may imagine then, that if a gas could be magnified billions of times, we should see what would look like a golfer's nightmare—an inferno of tiny, swiftly-moving, colliding molecules.

Liquids and solids.

Molecules attract each other with a force which is greater the nearer they are together. In a gas the molecules are too far apart for the attraction to be appreciable.

But when a gas condenses to a liquid, since a liquid has a definite and much smaller volume, the molecules must pack so closely together that they cannot get away from each other's attraction except by the comparatively slow process of evaporation.

On solidification, the molecules are packed even more closely together, since a solid has a definite shape. The cuttings on ancient gems are still sharp, proving that the molecules have not moved from their mean positions during thousands of years. Nevertheless, we believe that the molecules are still vibrating. We can imagine that, if it were possible to magnify a lump of solid matter millions of times, we should perceive that it was in a state of incessant quivering vibration, known as thermal agitation.

Thus the mutual attractions of the molecules in a solid are sufficient to prevent its disintegration as a result of thermal agitation (except, slightly, in the case of some substances such as camphor and iodine which change direct from solid to gas). However, when the temperature is raised the thermal agitation becomes more and more violent until melting results, then evaporation and finally boiling.

SUMMARY

The volume of a fixed mass of gas depends on its temperature and pressure.

Boyle's Law. The pressure and volume of a fixed mass of gas are inversely proportional, provided that the temperature is kept constant.

Charles' Law. All gases expand by $\frac{1}{273}$ of their volume at 0°C. for each 1°C. rise in temperature, provided that they are maintained at constant pressure.

Law of Pressures. The pressure of all gases increases by $\frac{1}{273}$ of their pressure at 0°C. for each 1°C. rise in temperature, provided that their volume is kept constant.

The lowest attainable temperature is -273°C. and is known as the **Absolute Zero.**

The Absolute Scale of temperature has for its zero -273°C. and rises in Centigrade degrees.

Thus $-273^{\circ}\text{C.} = 0^{\circ}\text{A.}, \quad t^{\circ}\text{C.} = 273 + t^{\circ}\text{A.}$

The constant-volume gas thermometer measures temperature by a determination of the pressure of the gas in the instrument at the given temperature.

The Kinetic Theory of Gases assumes a gas to consist of tiny moving particles, called molecules, with large empty spaces between them. The theory gives a satisfactory explanation of the gas laws, the absolute zero, and may be extended to the case of liquids and solids.

QUESTIONS

1. State Boyle's Law. Describe an experiment by means of which you could test its truth and explain fully how you would verify the law from your readings.

2. In order to investigate the relation between the pressure and volume of a gas at pressures less than atmospheric, Boyle used the

apparatus shown in Fig. 37. If the pressure of the atmosphere is 76 cm. of mercury, what is the pressure of the air in the tube?

Suppose the tube is raised until the height of the mercury in the tube is 60 cm. above that in the cistern. What will be the length of the air column?

3. The mercury in a faulty barometer stands at a height of 71.0 cm. above the reservoir and 10.0 cm. below the top of the tube. When the tube is lowered an extra distance of 12.0 cm. into the reservoir the mercury is found to stand at a height of only 65.0 above the reservoir. Calculate the height of the barometer. (1st M.B.)

4. A capillary tube sealed at one end and containing a mercury thread 15.8 cm. long is placed vertically (a) with the open end up, (b) with the open end down. The lengths of the air column between the mercury thread and the sealed end under these conditions are 14.7 cm. and 22.4 cm. respectively. Find the pressure of the atmosphere.

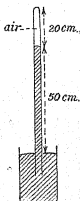


Fig. 37

5. State Charles' Law. Describe, with a diagram, how you would verify it experimentally.

6. State fully how the existence of an absolute zero of temperature (or the lowest temperature possible) was predicted. Why should there be such a temperature, and how far is the prediction borne out by experience?

7. To what temperature must a gas at $0^{\circ}\text{C}.$ be raised in order to double its volume (the pressure remaining constant)?

8. A column of air is trapped in a capillary tube sealed at one end, beneath a mercury index. The lengths of the air column at $19.5^{\circ}\text{C}.$ and $100^{\circ}\text{C}.$ are 14.7 cm. and 18.6 cm. respectively. Calculate the coefficient of expansion at constant pressure of the air. If this value were correct what would be the temperature of absolute zero?

9. Assuming that the density of air is 1.20 gm. per litre at $17^{\circ}\text{C}.$, what is its density at $100^{\circ}\text{C}.$? (C.)

10. Name and state the laws connecting (a) volume and temperature, (b) volume and pressure, in a gas. Deduce from them, showing clearly the steps in your proof, the law connecting the pressure and the temperature at constant volume. How would you test this law?

A glass vessel contains air at $60^{\circ}\text{C}.$ To what temperature must it be heated to expel one-third of the air, the pressure remaining constant? (L.)

11. Describe some method for determining the relation between the pressure and the temperature of a given mass of gas at constant volume.

A gas cylinder contains oxygen at a pressure of 14 atmospheres when the temperature is 7°C . By how much would the pressure increase if the temperature were raised to 37°C .? (C.)

12. State the laws which are found to hold for the relations between the pressure, volume and temperature of a given mass of a gas.

Air is pumped into a motor tyre until the pressure is 80 lb. wt. per sq. in. The air is then at a temperature of 17°C . Later in the day the temperature in the tyre rises to 57°C . Assuming that the tube does not stretch, what will then be the pressure? (O. & C.)

13. Describe a constant-volume air thermometer. The following readings were obtained with a constant-volume air thermometer on a day when the barometer stood at 760 mm.:

	Open limit (cm.)	Closed limit (cm.)
Bulb in melting ice	23.2	18.0
„ steam	53.6	18.0
„ solid carbon dioxide	3.2	18.0

Find the temperature of the solid carbon dioxide. (O. & C.)

14. When the bulb of a constant-volume air thermometer is immersed in ice, the mercury level in the open tube is 15 cm. higher than the fixed index. When the bulb is immersed in boiling water the difference in the two levels increases to 48.3 cm. When the bulb is immersed in a hot liquid the difference in levels decreases to 21 cm. What is the temperature of the hot liquid assuming that the atmospheric pressure is 76 cm. throughout? (O. & C.)

15. A spherical glass bottle whose capacity is 1000 c.c. at 0°C . is left for a time unstoppered in a room where the temperature is 15°C . and the mercury barometer stands at 755 mm. The stopper is inserted and the bottle is then immersed in a cooling mixture at -10°C . Find the pressure of the air in the bottle, assuming that no air passes the stopper during cooling. (The contraction of the glass may be neglected.) (N.)

16. The pressure in a constant-volume air thermometer is 770 mm. of mercury at 15°C .; what will it be at 20°C .? (O.)

17. How does the volume of a given mass of gas vary when its pressure and temperature are simultaneously changed?

A man requires to breathe the same mass of air per minute at the



top of a mountain as at the foot. The respective pressures and temperatures are 72 cm. and 17° C. at the foot and 48 cm. and 2° C. at the top. If he breathes 15 times per minute at the foot how often must he breathe at the top, assuming that he always fills his lungs to the same extent? (O. & C.)

18. Hot air at 700° C. and 30 atmospheres pressure is admitted to an engine. It is expanded to five times its volume and cooled to 200° C. At what pressure will it now leave the engine?

19. At the beginning of the compression stroke of a gas-engine the gas occupies 42 cu. in. at a pressure of 15 lb. per sq. in. and a temperature of 27° C. At the end of the compression stroke when the volume is 6 cu. in. the pressure is found to be 225 lb. per sq. in. What is the temperature of the gas? (C.)

20. Write a short account of the kinetic theory of gases. How does the theory explain (a) Boyle's Law, (b) the increase of pressure of a gas when its temperature is raised and its volume is kept constant, (c) the absolute zero?

Chapter IV

THE MEASUREMENT OF HEAT

When a beaker of water is heated by a bunsen burner, the temperature of the water rises. The cause of the rise in temperature is the *heat* which is being transferred from the bunsen burner to the water. We shall now consider how this heat may be measured.

Experiments with an electric heater.

Heat 500 gm. of water in a glass beaker with an electric heater and note the temperature at the end of each minute. In an experiment with a shaving-water heater the rise in temperature was found to be steady—about 6.2°C . per minute. We may assume that the heater was giving out heat uniformly and hence a uniform absorption of heat by the water resulted in a uniform rise of temperature.

When the experiment was repeated with the same heater using half the weight of water, 250 gm., the rate of rise of temperature was found to be approximately doubled, i.e. about 12.4°C . per minute.

It will be seen that

$$\begin{array}{rcccl} \text{Weight of water} \times \text{rise in temperature} & = & \text{constant,} \\ 500 \text{ gm.} & \times & 6.2^{\circ}\text{C.} & = & 250 \times 12.4. \end{array}$$

This suggests that the heat generated by the heater may actually be measured by the product of the weight of the water and the rise in temperature. In order to measure heat, however, we must first devise a unit.

The calorie.

The unit of heat used in scientific work is the calorie.

1 calorie is the quantity of heat required to raise the temperature of 1 gm. of water 1°C .

(We take 1 gm. rather than 1 c.c. of water since weight is unaffected by a rise in temperature but volume changes.)

Thus temperature is measured in degrees: heat is measured in calories.

Example. How many calories are required to raise 100 gm. of water from 15° C. to 27° C.?

1 gm. of water heated through	1° C. requires	1 calorie,
100 " "	1° C. require	100 calories,
100 " "	(27-15)° C. "	100 × 12
		= 1200 calories.

The difference between heat and temperature.

We have seen that when heat is put into a body its temperature rises, just as when water is put into a vessel the level rises. On this analogy, heat corresponds to quantity of water and temperature to level. It is so important to obtain a clear grasp of the difference between heat and temperature that we will pursue this analogy further.

Pour equal volumes of water into two glass cylinders of different cross-section. The levels in the two cylinders will be different. In a similar way if equal quantities of heat are communicated to a sewing needle and a steel crowbar by heating each with a lighted match, their temperatures rise by different amounts—one may well become red hot while the other remains cool to the touch.

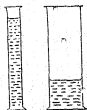


Fig. 38

Now if the two cylinders in Fig. 38 were connected by a tube at their lower ends, water would flow from the narrower to the wider cylinder until the levels were the same in each. Similarly heat flows from a body at a higher temperature to one at a lower temperature, until their temperatures are equal.

Again, a quantity of water is measured by pouring it into a measuring cylinder and finding the reading of its level. Similarly a quantity of heat is measured by putting it into a known mass of water and finding the rise in temperature with a thermometer.

Determination of the calorific value of coal by means of Darling's Fuel Calorimeter.

The calorific value of a fuel is the number of calories of heat given out when 1 gm. of it is completely burnt. It is of considerable importance to be able to measure this accurately. By this means, for example, good coal can be differentiated from bad, and physiologists can determine the energy value of various foods.

A calorimeter, as its name implies, is a vessel for measuring heat. The Darling Fuel Calorimeter (see Fig. 39) consists of an enclosed glass vessel, *A*, in which a measured quantity of coal is burnt, surrounded by water contained in a large glass cylinder, *B*. The hot gases produced by the combustion bubble through this water, and the heat which they impart to the water, is measured.

About 1 gm. of finely powdered coal is weighed in the nickel crucible *C* and placed in position with a fine wire dipping into it. The wire is connected to an electric battery via *D, D*, and heated by a momentary electric current to red heat, when it ignites the fuel. A gentle stream of oxygen is blown down the tube *E* from an oxygen cylinder and this enables the coal to burn easily and completely and at the same time prevents water from entering the vessel *A* through the holes in the plate *F*. The hot gases produced by the combustion pass down the hollow tube and bubble up through the holes in *F*, giving up their heat to the water. The temperature of the water is read at the beginning and end of the experiment. The weight of the water must be known. The glass cylinder and the vessel *A* will also take in heat, and this must be allowed for. The weight of water to which they are equivalent in this capacity is known as their *water equivalent*.

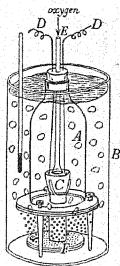


Fig. 39

The secret of accuracy in this experiment is to wait for about ten minutes after the coal is all burnt before taking the final temperature since it takes a long time for all the heat to be communicated to the water.

The heat emitted by the coal during combustion = (Weight of water + Water equivalent of apparatus) \times Rise in temperature.

British Units.

The calorie is not the only unit of heat employed. The *British Thermal Unit* (B.Th.U.) used by British engineers is the amount of heat required to raise the temperature of 1 lb. of water 1° F.

Gas is sold by its heat-producing properties, at so much a therm, 1 therm being 100,000 B.Th.U. A gas meter measures the

volume of gas passing through it in cu. ft. Thus the calorific value of the gas, in cu. ft. per therm, must be known in order to calculate its cost.

Specific heat.

It is found that if equal quantities of heat are put into equal weights of different substances the temperatures of the latter rise by differing amounts; water for instance requires more heat to raise its temperature than any other common substance, and liquids, on the whole, require far more heat than solids. We will compare the heat capacity of turpentine and water, using our electric heater.

Take a beaker, pour into it 250 gm. of turpentine, and heat for 1 min. with the electric heater. Repeat, using 250 gm. of water. It was found in an experiment that

Rise of temperature of turpentine = 26.4°C .

„ „ „ water = 12.4°C .

We can now calculate the number of calories of heat required to raise 1 gm. of turpentine through 1°C . 1 gm. of water, of course, requires 1 calorie, but 1 gm. of turpentine must require less than 0.5 calorie, since its rise in temperature is more than twice that of water.

Heat required to raise 250 gm. of water 12.4°C . = $250 \times 12.4 = 3100$ cal. Hence the heater must be giving out 3100 cal. per minute and this amount of heat must have been absorbed by the turpentine.

\therefore 250 gm. of turpentine raised 26.4°C . require 3100 cal.

1 gm. „ „ „ 1°C . requires $\frac{3100}{25 \times 26.4}$
= 0.47 cal.

This is known as the specific heat of turpentine.

Definition. The specific heat of a substance is the number of calories of heat required to raise 1 gm. of it through 1°C .

Example. A lump of copper, of 100 gm., is raised from 15°C . to 27°C . How much heat has been absorbed, given that the specific heat of copper is 0.1?

1 gm. of copper heated through	1° C. requires 0.1 cal.
100 gm. " "	1° C. require 0.1 × 100 cals.
100 gm. " "	(27-15)° C. "
	0.1 × 100 × 12 cals.
	= 120 cals.

Table of Specific Heats (cals. per gm. per ° C.).

<i>Solids</i>		<i>Liquids</i>	
Aluminium	0.21	Water	1.00
Iron	0.112	Meths.	.60
Copper	0.09	Paraffin	0.52
Brass	0.094	Turpentine	0.45
Mercury	0.034	Glycerine	0.58
Lead	0.032		
Glass	0.188-0.198		
Ice	0.504		
Earth, rock, sand	about 0.2		

The values of specific heats when expressed in B.Th.U. per lb. per ° F. are numerically the same as the above. Why?

The effect of the high specific heat of water on climate.

Look at the table of specific heats given above. It will be seen that the specific heat of water is five times as great as that of earth or sand. Consequently on a hot summer's day the temperature of dry sand rises much more rapidly than that of the sea; the sand feels hot to the touch, whereas the sea remains cool.

The temperature of the sea rises slowly since it absorbs a great deal of heat in the process: similarly its temperature falls slowly. This is the cause of the equable climate enjoyed by islands such as the British Isles. In the summer the sea keeps the island cool, and in the winter, warm. The average annual range of temperature at Lerwick in the Shetland Isles is 15° F., whereas at Olekminsk, on the same latitude in Siberia, the annual range is 95° F.

EXPERIMENTS IN CALORIMETRY.

Calorimetry is the study of the measurement of heat, and a calorimeter is a vessel, usually made of copper, employed for this purpose.

In the following experiments two types of measurement are involved: (i) readings of temperature, and (ii) weighings. The rise of temperature will, as a rule, be of the order of 10° C. or 20° C. Hence, in order to achieve an accuracy of 1 per cent., great care must be taken

to estimate each temperature to 0.1°C . Weights, on the other hand, will be of the order of 100 gm.: they need be taken only to the nearest gm. for the same degree of accuracy. Hence compression spring balances may be used, and by this means much time saved.

EXPERIMENT 1. *To find the heat generated by a bunsen burner per minute.*

Weigh a large copper calorimeter empty and three-quarters full of water. Read the temperature of the water. Heat the calorimeter and its contents with the bunsen for a measured time, so that its temperature rises about 20°C . Read the final temperature. Since 1 gm. of water requires 1 cal. to raise its temperature 1°C . and 1 gm. of copper requires 0.1 cal. (the specific heat of copper),

Heat taken in by the water

$$= 1 \times \text{Weight of water} \times \text{Rise in temperature.}$$

Heat taken in by the calorimeter

$$= 0.1 \times \text{Weight of calorimeter} \times \text{Rise in temperature.}$$

Hence calculate the total quantity of heat taken in, and therefore generated per minute.

EXPERIMENT 2. *To find the water equivalent of a metal calorimeter.*

The water equivalent of a calorimeter is the weight of water whose temperature rises by the same amount, when given the same quantity of heat, as the calorimeter.

Weigh the calorimeter (it should be massive) and find its temperature. Heat some water in another vessel to a temperature between about 30°C . and 35°C . Read the temperature of the warm water and pour it into the calorimeter until the latter is about one-third full. Stir, giving the calorimeter about 1 min. to absorb heat from the warm water, and then take the temperature of the water. Reweigh to find the weight of the warm water.

The warm water gives up some of its heat to the calorimeter until both are at the same temperature. The assumption underlying the method is that when the exchange of heat takes place no loss occurs in the process.

A source of error will therefore suggest itself. The calorimeter may lose heat to its surroundings while attaining its final temperature. To minimise this error place the calorimeter on a bad conductor, such as felt, inside a larger vessel to protect it from draughts.

Now the number of calories of heat given up by the warm water is equal to the product of its weight in gm. and the fall in temperature.

If x gm. is the water equivalent of the calorimeter, then the product of x and the rise in temperature of the calorimeter is equal to the heat gained by the calorimeter.

By equating the heat lost to the heat gained, x may be found.

You can check your result if the specific heat of the metal of which the calorimeter is composed is known. For

Water equivalent of calorimeter = Weight of calorimeter \times Specific Heat.

The reader should prove this relation for himself.

Another term, with a similar meaning, is sometimes used instead of water equivalent.

The thermal capacity of a calorimeter is the amount of heat required to raise its temperature by 1° .

Since 1 gm. of water requires 1 calorie to raise its temperature through 1° C., if the water equivalent of a calorimeter is 20 gm., its thermal capacity is 20 cal. per $^{\circ}$ C.

EXPERIMENT 3. *To find the specific heat of a metal by the method of mixtures.*

Weigh a lump of the metal, attach to it a piece of thread for lifting purposes, and heat it in boiling water. Take the temperature of the water.

Weigh a calorimeter empty and about two-thirds full of water. Take the temperature of the water.

- Transfer the hot solid by means of the thread rapidly to the calorimeter, stir and read the highest temperature reached.

The solid gives up some of its heat to the water and calorimeter, until all are at the same temperature. The heat received by the water and the calorimeter can be calculated and it is assumed that this is equal to the heat given up by the solid. From the value thus obtained of the heat given up by the known mass of the solid cooling through a known range of temperature, its specific heat can be found.

- As in Exp. 2 the calorimeter should be placed on felt inside a larger vessel, to minimise loss of heat to the surroundings during the experiment.

Again, the hot solid may lose heat while being transferred to the calorimeter. Also it is bound to carry over with it a little hot water for which no allowance is made. The best way to remedy these errors is to heat the solid in a special heater where it does not come in contact with the water, and from which it can be transferred to the calorimeter as quickly and directly as possible. Such a heater, however, is not always available.

The following readings were obtained using aluminium:

Readings.

Weight of solid	= 43.7 gm.
" " calorimeter	= 55.9
" " " + water	= 133.8 gm.
\therefore Weight of water	= 77.9
Initial temperature of water	= 15.3° C.
" " " solid	= 99.8° C.
\therefore Final temperature of solid and water	= 23.6° C.

Calculation.

Specific heat of copper	= 0.1.
\therefore Water equivalent of calorimeter	= 0.1×55.9
	= 5.6 gm.
Heat gained by water and calorimeter	= $(77.9 + 5.6) (23.6 - 15.3)$
	= 693 cal.
Let s = specific heat of solid.	
Heat lost by solid	= $43.7 (99.8 - 23.6) s$
	= 3330s cal.

But heat lost = heat gained,

i.e. $3330s = 693.$

$$\therefore s = \frac{693}{3330} \\ = 0.21 \text{ cal. per gm. per } ^\circ \text{C.}$$

EXPERIMENT 4. *To find the specific heat of a liquid by the method of mixtures.*

Transfer a solid of known specific heat from boiling water to a known mass of the liquid in a weighed calorimeter. Find the rise in temperature.

Take all the precautions against errors suggested in Exp. 3.

Find the heat lost by the solid and also the heat gained by the calorimeter and liquid in terms of s , the specific heat of the liquid. Form an equation representing

$$\text{Heat lost} = \text{Heat gained,}$$

and hence calculate s .

EXPERIMENT 5. *To determine the specific heat of a liquid by the method of cooling.*

The specific heat of a liquid may be determined by finding the time it takes to cool through a certain range of temperature and comparing it with the time taken by an approximately equal volume of water to cool through the same range of temperature under identical conditions.

The rates at which the two liquids lose *heat* to their surroundings are equal since they are cooling under the same conditions, but owing to the fact that their specific heats are different the rates at which their *temperatures* fall will *not* be equal.

Weigh a small calorimeter empty and about two-thirds full of the liquid. Warm the calorimeter until the temperature is slightly above 50°C . Place the calorimeter on felt, fit into it a cork through which pass a thermometer and a stirrer. (The cork is to prevent loss of heat by evaporation which would be different, probably, in the two cases.)

Stir steadily and find the time taken by the liquid to cool from, say, 50°C . to 45°C .

Repeat the experiment using the same calorimeter containing approximately the same volume of water. Why? It is essential that the water should be cooled through the same range of temperature, and that the temperature of the surroundings should be the same as before. Steady stirring is necessary so that the temperature of the whole bulk of the liquid is substantially the same as that of the outside of the calorimeter.

The following readings were obtained using glycerine:

	Glycerine	Water
Weight of empty calorimeter and stirrer =	67.6 gm.	67.6 gm.
„ liquid + „ „ =	141.4 gm.	122.7 gm.
„ liquid alone =	73.8 gm.	55.1 gm.
Time taken to cool from 50°C . to 45°C . =	369 sec.	466 sec.

$$\text{Water equivalent of calorimeter} = 67.6 \times .095 = 6.4 \text{ gm.}$$

Let s = specific heat of glycerine.

Heat lost by glycerine and calorimeter per sec.

$$= \frac{(6.4 + 73.8s)(50 - 45)}{369} \text{ cal.}$$

Heat lost by water and calorimeter per sec.

$$= \frac{(6.4 + 55.1)(50 - 45)}{466} \text{ cal.}$$

But the rates of cooling are equal,

$$\therefore \frac{(6.4 + 73.8s) 5}{369} = \frac{61.5 \times 5}{466}.$$

Hence

$$s = 0.57 \text{ cal. per gm. per } ^{\circ}\text{C}.$$

SUMMARY

Temperature is measured in degrees: heat is measured in calories of British Thermal Units.

1 calorie is the amount of heat required to raise 1 gm. of water through 1°C .

1 B.Th.U. is the amount of heat required to raise 1 lb. of water through 1°F .

The specific heat of a substance is the number of calories of heat required to raise 1 gm. of the substance through 1°C . (The same value for the specific heat is obtained if 1 lb. of water is raised through 1°F . and the heat is measured in B.Th.U.)

Heat absorbed by a body $= ms\theta$, where m = mass of body, s = specific heat, θ = rise in temperature.

The water equivalent of a body is the weight of water whose temperature rises by the same amount when given the same quantity of heat as the body.

Water equivalent of a body = Weight of body \times Specific heat.

The thermal capacity of a body is the amount of heat required to raise its temperature by 1° .

Determination of specific heats.

The specific heats of solids and liquids may be found by the method of mixtures. The specific heat of liquids may also be found by the method of cooling.

QUESTIONS

1. Define calorie, specific heat.

Using the table of specific heats on p. 57 find how much heat is required to raise

- (a) 150 gm. of water from 15°C . to 25°C .;
- (b) 100 gm. of copper from 15°C . to 25°C .;
- (c) 80 gm. of metal of specific heat s from 25°C . to 100°C .;
- (d) 10 lb. of iron from 32°F . to 212°F . (in B.Th.U.).

2. 200 gm. of water at 40°C . are mixed with 120 gm. of water at 16°C . What is the resulting temperature of the mixture?

3. Define water equivalent of a calorimeter.

60 gm. of water at 40.0°C . are poured into a calorimeter of mass 500 gm. and temperature 15.0°C . The resulting temperature is 28.0°C . Find the water equivalent of the calorimeter and the specific heat of the metal of which it is made.

4. (a) A lump of metal of mass 100 gm. is heated to 100°C . and then dropped into 150 gm. of water at 16.0°C . The resulting temperature is 26.0°C . Find the specific heat of the metal.

(b) If the water had been contained in a copper calorimeter of mass 80 gm. recalculate the specific heat of the metal allowing for the heat absorbed by the calorimeter. (Sp. ht. of copper = 0.1.)

5. Explain carefully:

(a) In December the weather is often warm in mid-Atlantic but colder on approaching the land.

(b) The human body can stand hot air without being burned, when hot water of the same temperature would be dangerous.

6. Does a body at 200°C . contain twice as much heat as it does at 100°C .? Give reasons.

7. A calorimeter of mass 50 gm. and specific heat 0.1 contains 200 gm. of water at 10°C . Determine the final resulting temperature when 100 gm. of water at 50°C . are poured into the calorimeter. (L.)

8. How much would 1 quart ($2\frac{1}{2}$ lb.) of water boiling at normal atmospheric pressure be cooled by pouring it into a silver vessel weighing 1 lb. at 60°F .? (Sp. ht. of silver = 0.06.) (L.)

9. Define British Thermal Unit; calorie.

What is the least cost of raising the temperature of 30 gallons of bath water through 45°F . by a gas heater if the bath has a water equivalent of 2 gallons and the price of gas is 9d. per therm? (1 therm = 100,000 B.Th.U.; 1 gallon of water weighs 10 lb.) (L.)

10. Two hot-water bottles of the same size and material are filled, the one with water and the other with mercury, at the same temperature. Which will remain warm the longer, and why? (The specific heat of mercury is 0.033 and its specific gravity is 13.5.) (C.)

11. What do you understand by the *thermal capacity* of a body?

0.4 gm. of a slow-burning powder is ignited in a closed copper vessel weighing 120 gm. completely immersed in 240 gm. of water contained in a calorimeter of water equivalent 48. If the original temperature is 13°C . and the resulting temperature is 42°C ., what is the heat developed on burning 1 gm. of powder?

If you were performing this experiment what precautions would you take to prevent gain or loss of heat by the calorimeter from the surroundings? (Sp. ht. of copper = 0.1.) (O.)

12. Explain the difference in the meanings of temperature and quantity of heat. On dropping 100 gm. of steel at 315°C . upon 2500 gm. of mercury at 15°C . the temperature of the liquid rose to 51°C . The specific heat of mercury is 0.033; what was that of the steel? (N.)

13. A piece of iron weighing 100 gm. is heated for some time in a bunsen flame. It is then plunged into 250 gm. of water, and causes a rise of temperature from 12.0°C . to 30.0°C . Find the temperatures of the bunsen flame. (Sp. ht. of iron = 0.11.)

14. How do you find the water equivalent of a calorimeter in the school laboratory? What use have you known to be made of the information so obtained? (N.)

15. How would you find the temperature of a red-hot iron ball if only mercury thermometers were available? (O.)

16. Describe the precautions taken to prevent loss of heat from a calorimeter. What is meant by the water equivalent?

The following half-minute readings of a calorimeter were made, a hot piece of metal being dropped in at the end of the second half-minute:

Time in $\frac{1}{2}$ min.	0	1	2	3	4	5	6
Temperature $^{\circ}\text{C}$.	15.10	15.00	Metal dropped in	18.97	18.45	18.00	17.60

Explain with a graph exactly how you would estimate the temperature rise, giving reasons. (O.)

17. Explain what is meant by the specific heat and by the water equivalent of a body.

An aluminium calorimeter weighs 28.40 gm. empty, and 80.70 gm. after some water has been poured into it. An aluminium cylinder weighing 79.60 gm. is heated to 98.8°C ., and dropped into the water. The temperature of the water rises from 15.2°C . to 34.0°C . Calculate the specific heat of the aluminium and the water equivalent of the calorimeter. (C.)

18. A liquid of specific heat 0.5 and temperature 30°C . is mixed with another liquid of specific heat 0.4 and temperature 15°C ., and the final temperature is found to be 20°C . In what proportions by weight were the liquids mixed? (C.)

19. Define specific heat.

A concert hall contains 1000 people and requires for efficient ventilation 8 litres of air per sec. per person. This air has to be supplied at 15°C . while the outside temperature is 2°C . Calculate the heat required during a 2-hour performance. (Sp. ht. of air = 0.24 ; density of air = 1.3 gm. per litre.) (O. & C.)

20. A silver teapot weighs 300 gm. Tea is made in it from 20 gm. of tea leaves and 600 gm. of water at 100°C . If the temperature of the room is 15°C ., what difference will it make to the initial temperature of the tea if the pot is first heated to 80°C .? (Sp. ht. of silver = 0.056 ; sp. ht. of tea leaves = 0.5 .) (O. & C.)

21. A student gives up the following set of readings for the determination of the specific heat of iron:

Weight of calorimeter	= 56 gm.
" " water	= 173.215 gm.
" " iron	= 2.9 gm.
1st temperature of water	= 16°C .
2nd " "	= $16\frac{1}{4}^{\circ}\text{C}$.
1st " iron	= 212.37°F .

Enumerate the errors made by the student in taking his readings.

22. A copper calorimeter weighing 20 gm. and containing 63 gm. of glycerine cooled from 50°C . to 40°C . in 196 sec. An equal volume of water, weighing 50 gm., and cooling under the same conditions, took 420 sec. to cool from 50°C . to 40°C . Calculate the specific heat of glycerine. (Sp. ht. of copper = 0.1 .) (O. & C.)

23. A calorimeter of mass 100 gm. and specific heat 0.1 containing 150 gm. of water cools from 60°C . to 55°C . in 4 min. If the same calorimeter containing an equal volume of glycerine of specific heat 0.6 and specific gravity 1.25 is allowed to cool through the same range, how long will it take? (O. & C.)

Chapter V

CHANGE OF STATE

In this chapter we shall consider the heat absorbed when substances change their state from solid to liquid and liquid to vapour, and also the temperatures at which these changes take place, namely the melting and boiling points. Apart from its intrinsic theoretical interest this branch of heat has important practical applications in the refrigerator and steam engine.

LATENT HEAT

Take a light copper calorimeter and fill it with small pieces of ice dried quickly and carefully with blotting paper. Heat it with a small steady bunsen flame and stir with a thermometer. The temperature remains at 0°C . until all the ice has melted. Can you think of any reason why this should be so? Note the time when the heating commenced and also when all the ice has melted. Continue the heating and note the time when the water reaches its boiling point. See if you can calculate from these readings of temperature how many calories have to be imparted to 1 gm. of the ice before its temperature can be raised above 0°C . This quantity is known as the Latent Heat of fusion of ice, and is the heat required to change the state of 1 gm. of ice to water without change of temperature.

The following is a set of readings:

Time required to melt ice = 173 sec.

Time required to raise the water from 0°C . to 100°C .
= 276 sec.

Consider 1 gm. of the ice (or water).

To raise 1 gm. of water from 0°C . to 100°C . requires 100 cal.

\therefore 1 gm. of water receives in 276 sec. 100 cal.,

1 gm. of ice	„	173	„	$100 \times \frac{173}{276}$
				= 63 cal.

This experiment is a rough one since no allowance is made for the heat absorbed by the calorimeter or for the heat gains and losses during the experiment from and to the surroundings. More

accurate experiments show that 80 cal. of heat are required to melt 1 gm. of ice without raising its temperature.

The phenomenon was first investigated by Joseph Black (1728-99), a professor at Glasgow University. He asked himself what happened to the heat absorbed by the ice when the latter was converted into water. Since it did not manifest itself by a rise in temperature, Black called this heat *latent* (or hidden) heat.

The heat absorbed when unit mass of ice is converted into water at the same temperature is called the latent heat of fusion of ice.

Its value is **80 cal. per gm.** or **144 B.Th.U. per lb.**

This latent heat does apparently disappear but it is not lost or destroyed, for when 1 gm. of water at 0°C . changes back into ice it gives out 80 cal. of heat. The mild winters experienced in the neighbourhood of the Great Lakes of North America are due to the large quantity of heat given out when the surfaces of these lakes freeze.

The phenomenon is not confined to ice but is exhibited by all substances. For example, 1 gm. of solid lead at 327°C . (its melting point) requires 5.36 cal. to convert it into molten lead at the same temperature.

EXPERIMENT. *To make an accurate determination of the latent heat of fusion of ice.*

Weigh a calorimeter empty and also rather more than half full of water. Heat it to about 25°C . or 30°C . Place it on felt inside a larger vessel and find its temperature. Dry some large pieces of ice with blotting paper, and drop them into the calorimeter. Stir, and when the ice has melted, read the temperature, which should not be allowed to fall more than one or two degrees below that of the room; otherwise dew may be deposited on the outside of the calorimeter. Now reweigh the calorimeter and its contents in order to find the mass of ice used.

The following is a set of readings:

Readings.

Weight of calorimeter	= 43.9 gm.
" " " + water	= 149.5 gm.
\therefore Weight of water	= 105.6 gm.
" " calorimeter + water + ice	= 165.3 gm.
\therefore Weight of ice	= 15.8 gm.
Initial temperature of calorimeter and contents	= 28.0°C .
Final " " " " " "	= 14.2°C .

Calculation.

$$\begin{aligned}
 \text{Water equivalent of calorimeter} &= 43.9 \times 0.1 \\
 &= 4.4 \text{ gm.} \\
 \text{Heat lost by water and calorimeter} &= (105.6 \times 4.4) (28.0 - 14.2) \\
 &= 1518 \text{ cal.} \\
 \text{Let } L \text{ cal. per gm.} & \\
 \text{Heat required to melt the ice} &= 15.8 L. \\
 \text{Heat required to raise the melted ice} & \\
 \text{from } 0^\circ \text{ C. to } 14.2^\circ \text{ C.} &= 15.8 \times 14.2 \\
 &= 224 \text{ cal.} \\
 \text{But} & \quad \text{Heat gained} = \text{Heat lost,} \\
 \text{i.e.} & \quad 15.8L + 224 = 1518. \\
 & \quad \therefore L = \frac{1294}{15.8} \\
 & \quad = 82 \text{ cal. per gm.}
 \end{aligned}$$

This experiment gives results with a consistency of between 3 and 4 per cent.

Black's ice calorimeter.

Once the value of the latent heat of fusion of ice has been accurately determined, it may be used to find the specific heat of a solid. The apparatus required is known as an ice calorimeter, and its first and simplest form, shown in Fig. 40, was devised by Black. A cavity is made in a large block of ice and closed by a lid of ice. This cavity is carefully dried. A solid of known weight is heated, say in boiling water, and then quickly transferred to the cavity. It will cool to 0° C. and the heat it gives out will all be used in converting some ice to water at 0° C. The weight of the ice melted may be determined by absorbing it with a weighed sponge and reweighing the sponge. The heat given out by the solid can then be calculated since each gm. of ice melted requires 80 cal. of heat.



Fig. 40

Latent heat of vaporisation of water.

If a can of water is heated by a bunsen burner its temperature will rise until its boiling point is reached. While the water boils away however no further rise of temperature takes place, and stronger heating merely causes the water to boil away more quickly. When boiling water is converted into steam, therefore,

there is an absorption of heat without rise in temperature similar to the case of ice on melting.

The latent heat of vaporisation of water is the heat required to change unit mass of water at its boiling point into steam at the same temperature. Its value is 536 cal. per gm. or 965 B.Th.U. per lb.

A rough determination of the latent heat of vaporisation may be made by heating a can of cold water with a steady bunsen flame and finding the times required for the water to reach the boiling point, and then to boil away completely. You should perform this experiment for yourself, calculating the result by a method similar to that used on p. 66.

When steam condenses to water it gives out its latent heat.

Example. 5 gm. of steam at 100°C . condense to water at 15°C . How much heat is given out?

$$\begin{aligned}\text{Heat given out when steam condenses to water at } 100^{\circ}\text{C.} &= 5 \times 536 \\ &= 2680 \text{ cal.} \\ \text{Heat given out when condensed water at } 100^{\circ}\text{C.} &\text{cools to } 15^{\circ}\text{C.} \\ &= 5(100 - 15) \\ &= 425 \text{ cal.} \\ \therefore \text{Total heat given out} &= 2680 + 425 = 3105 \text{ cal.}\end{aligned}$$

EXPERIMENT. *Accurate determination of the latent heat of vaporisation of water.*

Weigh a calorimeter empty and about three-quarters full of water. Read the temperature of the water. Blow in steam until the temperature rises to between 30°C . and 35°C . Stir and take the final temperature. Reweigh to find the mass of steam condensed. If the final temperature is higher than 35°C . the final weighing will not give the correct weight of condensed steam owing to rapid evaporation.

A special water trap must be fitted to the boiler generating the steam (See Fig. 41) to prevent drops of water entering the calorimeter with the steam. Indeed this is a serious source of error and one difficult to eliminate entirely. The end of the glass tube *A* should be placed under the surface of the water in the calorimeter.

To work out the result find:

- the heat gained by the water and calorimeter;
- the heat given up by the steam in condensing to water at 100°C . (using *x* for the unknown latent heat);

(c) the heat given up by the condensed steam at 100°C . cooling to the final temperature.

Then obtain an equation for x from the fact that $(a) = (b) + (c)$.

It is not easy to obtain accurate results in this experiment at the first attempt. The correct value is 536 cal. per gm. Black obtained the value 417 and later 450. However, the method is capable of an accuracy of about 5 per cent.

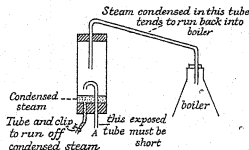


Fig. 41

Cooling by evaporation.

If water is left in a shallow dish for a considerable time it will eventually disappear. It is said to *evaporate*. Thus a liquid may change into vapour at temperatures other than its boiling point, although the process of evaporation is much slower than that of boiling. It takes place only at the surface, and not within the body of the liquid as in boiling.

When a liquid evaporates it must extract its latent heat of vaporisation from its surroundings. If petrol or ether is spilt on the hand one can feel heat being extracted from the hand as the liquid rapidly evaporates. The chief mechanism for cooling the human body is the excretion of perspiration from the pores of the skin, and the subsequent evaporation of water from the surface of the body. Normally we perspire from a pint to a quart of fluid in twenty-four hours: but in a game calling for vigorous exercise one may well perspire far more than this in an hour. It is not advisable, however, to stand about in wet clothing, for the latent heat of vaporisation which is taken from the body by this unnatural means may cause a severe chill. The Scotch shepherd, when his plaid is wet, is said to turn it inside out with the wet side innermost to lessen the rate of evaporation, and the risk of chill.

The following simple experiment is a further demonstration of cooling due to evaporation: Place a glass beaker on a block of wood on which one or two drops of water have been sprinkled. Pour some ether into the beaker and cause it to evaporate rapidly by blowing air through it. Soon the drops of water will freeze, and the beaker will become attached by means of the ice to the wood.

The melting point.

When a solid changes into a liquid it is said to melt: in the case of pure substances the change takes place always at a definite fixed temperature called the *melting point*. At this temperature solid and liquid may remain in contact without there being any tendency for one form to change into the other.

EXPERIMENT. *To obtain a cooling curve for molten lead as it solidifies, using a thermo-couple.*

Lead melts at 327°C. ; a special type of thermometer, called a thermo-couple (see p. 10) is used to record its temperature.

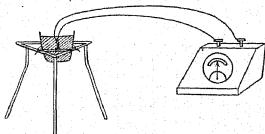


Fig. 42

Twist together the ends of two wires, one of copper and the other of eureka (an alloy composed of 60 per cent. copper and 40 per cent. nickel). Connect the two free ends to the terminals of a sensitive (Weston) galvanometer—an instrument for measuring electric currents. Your simple thermo-couple is now ready for use.

Melt some lead in a crucible. Place the junction of the wires in the lead (see Fig. 42). A deflection of the galvanometer needle will be produced. Heat the lead a little above its melting point, and remove the bunsen.

Read the deflection of the galvanometer at half-minute intervals for about 10 min. (i.e. until some time after the lead has solidified).

Plot time (horizontal axis) against galvanometer deflection which is a measure of the temperature, on graph paper. Your graph will look like that shown in Fig. 43.

It will be seen from the graph that the temperature of the molten lead falls steadily, until suddenly it becomes stationary ceases to fall. After some time the temperature begins to fall again.

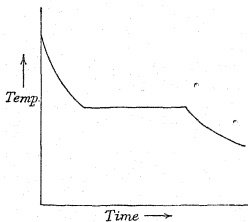


Fig. 43

How are we to account for the horizontal portion of the graph? There is no reason to suppose that the lead should suddenly cease to lose heat to its surroundings. The answer is that at this time the lead was solidifying and *as it solidified it gave up its latent heat thereby compensating for the loss of heat to the surroundings and arresting the fall of temperature.* Thus the reading of the galvanometer corresponding to the horizontal portion of the graph represents a temperature of 327°C. , the melting point of lead.

Determination of melting point.

EXPERIMENT 1. An experiment similar to the foregoing may be used to determine the melting point of a substance. Heat some naphthalene in a test-tube, preferably in boiling water rather than with a naked flame since naphthalene is very inflammable.

Place a thermometer in the liquid naphthalene, dry the outside of the test-tube, and read the temperature of the naphthalene every minute until some time after the latter has solidified. (A range of from 90°C. to 70°C. is suitable.)

Plot temperature against time, when a graph similar to that in Fig. 43 will be obtained. The temperature corresponding to the horizontal portion of the graph will give the melting point.

If the substance, like paraffin wax for instance, is a mixture, the various constituents will melt at different temperatures. The graph, when paraffin wax is solidifying, is not parallel to the time axis, but slopes. There is no sharp melting point in this case.

EXPERIMENT 2. A determination of melting point is often made in qualitative analysis as a further clue to the identity of the unknown substance. It often happens, however, that only a small quantity of the substance is available for the melting-point determination, and in this case the following method is suitable.

Heat some glass tubing in a bunsen flame, and draw it out into a capillary tube. Cut off a suitable length of capillary tube, seal one end, and put a little of the substance into it. Fasten the tube to the bulb of a thermometer by means of rubber bands. Heat in a water-bath and note the temperature when the substance melts.

Table of Melting Points.

Tungsten	3000° C.	Naphthalene	80° C.
Platinum	1775° C.	Wood's metal	70° C.
Glass	1000-1400° C.	Paraffin wax about	54° C.
Cast iron	1200° C.	Ice	0° C.
Copper	1083° C.	Mercury	- 39° C.
Lead	327° C.	Alcohol	- 112° C.
Tin	232° C.	Oxygen	- 219° C.

Change of volume on melting and solidification.

When water freezes it expands; 10 c.c. of water becomes 11 c.c. of ice. Hence ice has a smaller density than water and floats in water. Water which has no room to expand on freezing will exert an enormous force. This is the cause of burst water-pipes during a frost. It is also a cause of the breaking up of rocks and the weathering of soil. Water percolating into clefts in the rocks on a mountain side during the day splits off pieces of stone when it freezes during the night.

Most substances contract on solidifying. A striking example is paraffin wax. Melt some paraffin wax in a boiling tube. When it cools, the wax in contact with the surface of the boiling tube solidifies first and the contraction causes a deep crater to form in the surface of the wax.

Iron and type metal are two of the few metals which, like water, expand on solidification. They can be used for making castings, since on solidifying they fill the mould completely. Gold and silver coins, on the other hand, have to be stamped, they cannot be cast since these metals shrink on solidifying.

Effect of pressure on the melting (or freezing) point.

An increase of pressure lowers the freezing point of water. Thus ice subjected to a sufficient pressure will melt. For suppose the ice is at 0°C . and the pressure depresses the freezing point to -0.1°C ., the ice is now at a temperature above its freezing point, and will therefore melt. A pressure of 1 atmosphere depresses the freezing point of water by 0.0076°C .

If ice is subjected to a pressure and then the pressure is removed the ice will first melt and then freeze again. The phenomenon is known as *Regelation* (literally—re-freezing), and may be demonstrated by the following simple experiment.

Hang a heavy weight by means of fine copper wire over a block of ice (see Fig. 44). The wire will slowly cut through the ice but leave the block undivided. Since the area of the wire in contact with the ice is small it exerts a large pressure on the ice, depressing its freezing point. Hence the ice immediately below the wire melts but freezes again as the wire descends below it and the pressure is removed. The latent heat

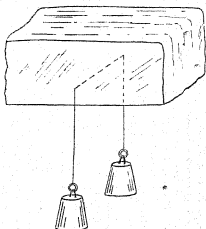


Fig. 44

evolved when the water above the wire freezes passes through the wire and melts the ice below. Hence a bad conductor of heat, such as string, will not cut through ice in this way.

When skating, the ice below the blade of the skate melts and the water acts as a lubricant. The pressure exerted by the skate lowers the freezing point of the water. If, however, the ice is very cold, the depression of the freezing point is not sufficient to make the ice melt. Under these circumstances the skate will not bite and skating is difficult.

The binding together of a snow-ball is a further example of regelation. When the snow-ball is compressed the particles of snow begin to melt and freeze again when the pressure is removed, forming a comparatively solid ball. On very cold days snow-balls cannot be made. Why?

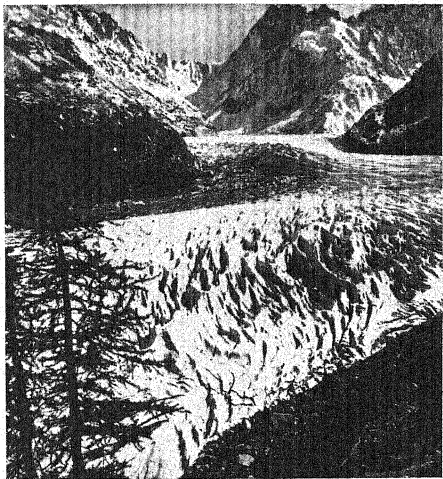


Fig. 45. Mer de Glace Glacier. The ice flows like a river at the rate of about a foot a day. This is an example of regelation.

Again, when walking on snow, a crust of hard snow forms under and sticks to the boot. Why is this?

Perhaps the most interesting example of regelation is the flow of a glacier, a river of ice. Glaciers flow down the high valleys of

great mountain ranges such as the Himalayas and the Alps at the rate of a foot or so a day. Tributary glaciers weld themselves into a trunk glacier, and each glacier narrows and widens to fit the shape of the valley through which it is flowing. The effects occur as a result of the slow melting point of the ice under pressure and the subsequent refreezing when the pressure is eased or removed.

The freezing point of iron, like that of water, is lowered by pressure; that of paraffin wax, however, is raised. The following law holds for all substances:

Pressure lowers the freezing points of substances which expand on solidifying, but raises the freezing points of substances which contract on solidifying.

Effect of impurity on the melting point.

When salt is added to ice the melting point is lowered. The effect is a large one depending on the concentration of the salt solution; the melting point may be depressed to as low a temperature as -22°C .

Consequently when salt is mixed with ice some of the ice melts, absorbing the necessary latent heat from the mixture until the temperature of the whole has dropped to the new melting point. This makes a very convenient freezing mixture.

The effect is a general one occurring with other substances besides ice and salt.

Vapour pressure.

Fill a barometer tube with mercury, place the thumb over the open end, invert, plunge the end closed by the thumb under the surface of mercury in a trough and remove the thumb. The mercury in the tube will stand to a height of about 76 cm. above the level of the mercury in the trough: the space above the mercury in the tube is a vacuum. Introduce water into this space by means of a piece of bent glass tubing (see Fig. 46), by blowing a drop of water to the surface of the water in the barometer tube. The mercury in the tube will drop slightly. This is due to the fact that the water, on reaching the space above the mercury, evaporates and the water vapour exerts a pressure, forcing the mercury down. Introduce more water above the mercury in the tube. After a time the water ceases to evaporate

and rests on the surface of the mercury. The space is said to be *saturated* with water vapour and the mercury cannot be depressed further. It is desirable to have two barometer tubes, one containing no water vapour. Then the difference in level in the two tubes, or the depression of the mercury which has taken place in the tube containing the water vapour, is called the **saturated vapour pressure** of water at the given temperature.

Every liquid has its own characteristic saturated vapour pressure at a particular temperature. It is instructive to have four barometer tubes side by side and compare the saturated vapour pressures of water, methylated spirits and ether. At 15°C . the values are water 1.28 cm., methylated spirits 7.2 cm., ether 35.8 cm. of mercury.

Another term, the *vapour pressure of a liquid*, is sometimes used to denote the pressure of its saturated vapour.

Evaporation is rapid when the space above a liquid contains very little of its vapour but becomes slower as the vapour approaches saturation. Every housewife knows that on a damp muggy day the washing dries more slowly than on a dry day.

The variation of saturated vapour pressure with temperature.

As the temperature rises the vapour pressure of a liquid increases.

The vapour pressure of water, for instance, at temperatures above that of the atmosphere may be found by surrounding the barometer tube used in the previous experiment with a water-jacket (see Fig. 47), heating with an electric heater and taking simultaneous readings of the temperature of the water and the level of the mercury. The saturated vapour pressure at each temperature is equal to the difference in the heights of the mercury barometer and the mercury in the tube containing the water vapour. Efficient stirring is essential and there should always be a thin layer of water resting on the mercury surface in the tube to ensure that the space above it is saturated.



Fig. 46

Fig. 48 is a graph showing the results obtained from such an experiment. It will be seen that the vapour pressure increases slowly with temperature at first, and then more rapidly. At 100°C ., the boiling point of water, the vapour pressure of the water is equal to the pressure of the atmosphere.

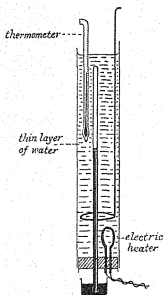


Fig. 47

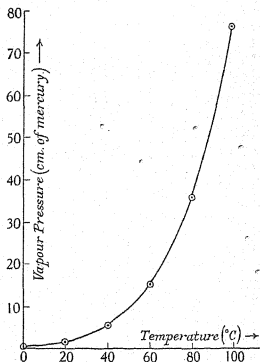


Fig. 48. Variation of the vapour pressure of water with temperature.

Results to a temperature as high as this, however, are not possible with this apparatus, since the mercury in the tube would be depressed to the same level as that in the trough and the water-jacket is not sufficiently long.

Dalton's law of partial pressures.

If a liquid is allowed to evaporate into air, at a pressure, say, of 76 cm. of mercury, instead of into a vacuum, the rate of evaporation will be slower but the final saturated vapour pressure in the air will be the same. Since the vapour pressure of water at 15°C . is 1.28 cm., the resulting total pressure when the air is saturated will be $76 + 1.28 = 77.28$ cm. of mercury.

This illustrates what is known as *Dalton's law of partial pressures*. If two gases (or vapours) occupy a space, each exerts the same pressure that it would exert in that space if the other were absent.

The difference between a gas and a vapour.

We speak of air as a gas, but of steam as a vapour. The main physical difference between them is that the latter is near its liquefying point but the former is not. It is natural to enquire whether, if air were cooled sufficiently, it would be correct to call it a vapour, and if steam were heated sufficiently it would be correct to call it a gas.

Now there is a temperature, called the critical temperature, differing for every gas, above which it is not possible to liquefy the gas by pressure alone. The critical temperature of steam is 365°C . and of air -140°C .

It is the custom to use the term 'vapour' for a substance below its critical temperature, and 'gas' when it is above that temperature.

Boiling.

Boiling differs from evaporation in that the liquid changes to vapour throughout its bulk instead of merely at the surface. The bubbles which rise to the surface as soon as the heating of cold water commences consist of gases dissolved in the water. When a kettle begins to sing bubbles of steam are formed in the water at the bottom nearest the flame, which has reached 100°C ., but on rising through the cooler water above they condense. On collapsing these bubbles create the familiar noise of "singing".

When, however, the whole liquid is at 100°C ., the bubbles of steam rise to the surface and enter the atmosphere.

Thus a liquid boils when its vapour pressure is equal to that of the atmosphere above it.

An experiment to show that, at its boiling point, the vapour pressure of a liquid is equal to that of the atmosphere.

Pour mercury into a glass U-tube, one arm of which is sealed, and by tilting drive out all the air from the sealed arm. Introduce a drop or two of water, which has been boiled to exclude all

dissolved air, into the sealed arm (see Fig. 49*a*). Place the U-tube in a flask (see Fig. 49*b*) and heat it with the steam from boiling water. The mercury in the sealed arm will be forced down until the levels in the two arms of the U-tube are at the same height, showing that the pressure of the water vapour in the sealed arm is equal to that of the atmosphere.

This method may be used to find the boiling point of a liquid of which a small quantity only is available. Introduce the liquid into a U-tube similar to that described above, heat in a water-bath, and find the temperature of the water-bath when the mercury levels in the U-tube are the same. This is equal to the boiling point of the liquid. If the liquid boils at a temperature higher than 100°C. , another liquid with a sufficiently high boiling point must be used for the bath.

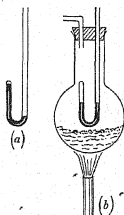


Fig. 49. The drop of water at the top of the left-hand sealed arm in (a) vaporises when heated and in (b), at the temperature of boiling water, its vapour pressure is equal to the pressure of the atmosphere.

The effect of pressure and dissolved impurities on the boiling point of a liquid.

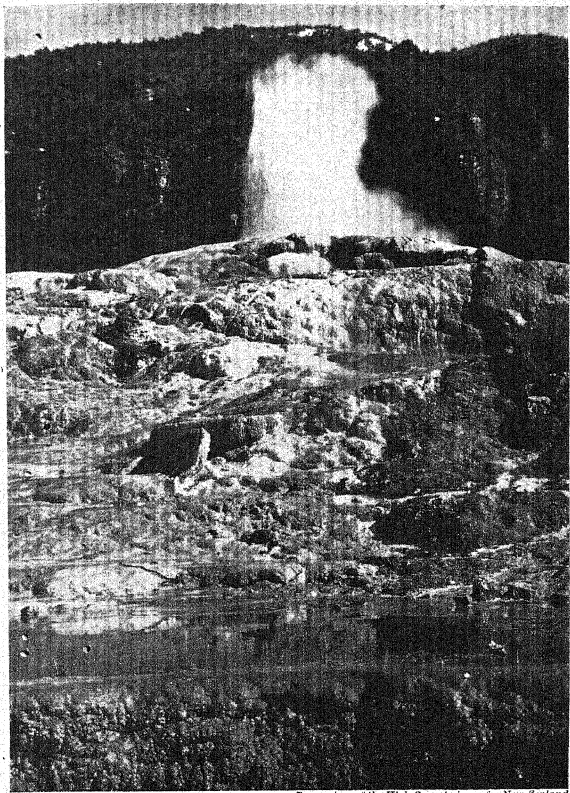
If the pressure of the atmosphere above a liquid is reduced the liquid need not be raised to so high a temperature for its vapour pressure to become equal to the pressure of the atmosphere. It therefore boils at a lower temperature. Similarly, increasing the pressure above a liquid raises its boiling point.

The boiling point of a solution is always higher than that of the pure solvent.

These facts may be verified in the school laboratory by the experiment described on p. 13.

Geysers.

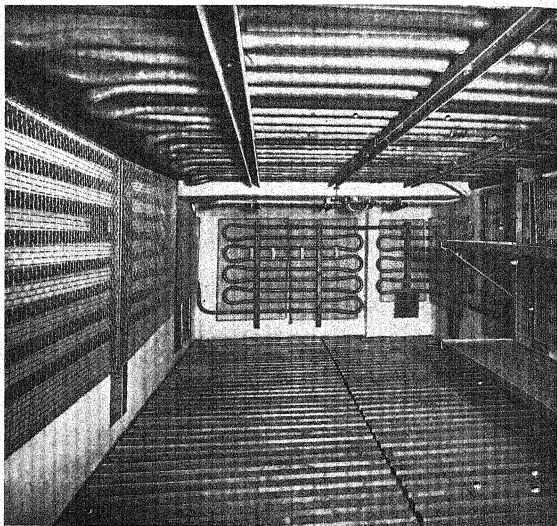
In the North Island of New Zealand there is a volcanic region where a number of geysers are to be found. These are holes in the ground from which, intermittently, boiling water is hurled into the air, sometimes to a height of 100 ft. (see Fig. 50).



By courtesy of the High Commissioner for New Zealand

Fig. 50. Waikite Geyser, North Island, New Zealand. A spout of steam and boiling water is thrown, at intervals, high into the air (see pp. 80-3).

The phenomenon is an interesting natural example of the raising of the boiling point of water by pressure. The holes are filled with water which is heated from below by steam generated by volcanoes. The water at the bottom is under the pressure of the water above it, and consequently, although its temperature may be over 120°C ., it is below its boiling point at that pressure.



By courtesy of the Liverpool Refrigeration Co. Ltd.

Fig. 51. A refrigerated chamber of a ship for the cold storage of meat. Note the pipes along the ceiling and walls: these carry brine cooled to a temperature below 0°C . by a central refrigerating plant. Note also the cement-faced insulation behind the pipes which prevents heat from entering the chamber from the outside.

Immediately prior to an eruption detonations are heard and the water level is raised, causing an overflow. This is due to boiling at the bottom. Owing to the overflow the pressure on the whole of the liquid is reduced and it is consequently much nearer to its boiling point. Soon it boils and as the pressure on it is removed there is great evolution of steam since it is well above its boiling point at normal atmospheric pressure. The steam throws a spout high into the air.

Refrigerators.

Instruments for abstracting heat are known as refrigerators.

Cold storage is widely used for keeping food fresh, since the germs which cause food to go bad are inactive in the cold. Above

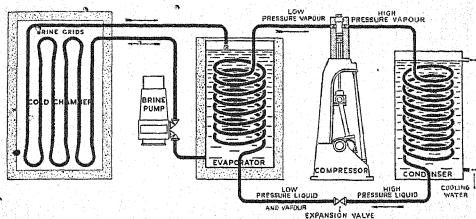


Fig. 52

50° F., for example, the germs which turn milk sour multiply rapidly: below this temperature they do not. Foods such as meat, butter and cheese, and bananas, are brought to this country from the Argentine, New Zealand, the Canaries, etc., in chambers on board ship similar to that shown in Fig. 51. The chamber is cooled by pipes through which circulates brine, kept at a temperature below 0° C. by a refrigerating plant. The pipes are attached to the ceiling rather than near or below the floor. The air near the pipes becomes colder and hence heavier than the air below it: it therefore falls, causing a continuous circulation.

The principle of the most common form of refrigerator is as follows. An easily vaporisable liquid such as ammonia, carbon dioxide, or sulphur dioxide is caused to evaporate by reducing the pressure above it. While evaporating it extracts its latent

heat of vaporisation from its surroundings—either the refrigerating chamber *or* brine which, circulating through pipes, can convey the refrigerating effect to a number of distant chambers (see Fig. 52).

The vapour is then compressed by a pump in order to liquefy it again. It now *gives out* its latent heat, which is removed by air



By courtesy of the British Thomson-Houston Co. Ltd.

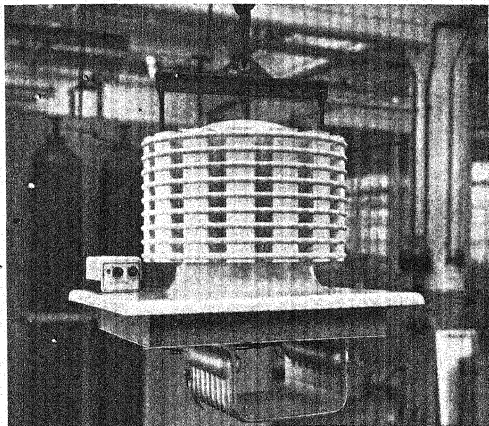
Fig. 53. A refrigerator for ordinary domestic use.

or water cooling. Once more the liquid is allowed to evaporate, taking in heat from its surroundings, and the cycle of operations is repeated.

Large cargo vessels carry refrigerating plant capable of cooling insulated chambers with a combined capacity of over half a million cubic feet. The ordinary domestic refrigerator with its capacity of a dozen or so cubic feet, works on the same principle, and is coming into wide use. Fig. 53 shows such a refrigerator made by the British Thomson-Houston Company, and Fig. 54

shows the complete compressor (and refrigerating) unit detached from the cabinet.

In Fig. 54 the hollow metal U-shaped vessel below the cabinet top is the evaporator or chilling unit. It contains liquid sulphur dioxide which is caused to evaporate there and hence cool the cabinet. To produce ice rapidly trays containing water are placed in the U (see the trays in Fig. 53).

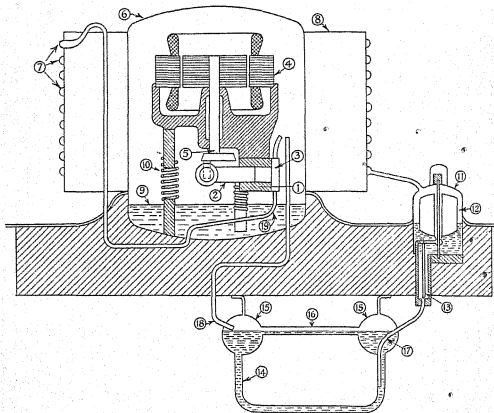


By courtesy of the British Thomson-Houston Co. Ltd.

Fig. 54. The compressor and refrigerating unit detached from the complete refrigerator shown in Fig. 53.

The upper part of the unit contains the compressor pump, and an electric motor to work it. Round the outside can be seen copper pipes wrapped round steel fins, which serve as an air-cooler for the sulphur dioxide when liquefied.

Fig. 55 is a diagram showing the interior of the unit. The evaporator or chilling unit (14) communicates with the compressor case (6) through a pipe (18). When the pump, (1) cylinder, (2) piston, (3) valve box, begins to suck in sulphur dioxide vapour,



By courtesy of the British Thomson-Houston Co. Ltd.

Fig. 55. Diagram of the interior of the compressor and refrigerating unit shown in Fig. 54. The key to the numbers is as follows: 1. Cylinder; 2. Piston; 3. Valve Box; 4. Motor; 5. Steel crank shaft; 6. Drawn steel case; 7. Copper condenser coils; 8. Steel fins; 9. Reservoir of lubricating oil; 10. Three-point spring support; 11. Float valve; 12. Hollow steel float; 13. Needle valve; 14. Evaporator, or chilling unit; 15. Cylindrical headers; 16. Equalizer tube; 17. Injector tube; 18. Suction pipe; 19. Tube through which compressed vapour passes to the condenser coils.

the pressure over the liquid sulphur dioxide in the evaporator gets less, and hence the liquid evaporates, taking its latent heat from the cabinet of the refrigerator. (Incidentally the liquid

marked (9) is lubricating oil, which is supplied to all moving parts by a pump not shown in the diagram.)

The pump compresses the sulphur dioxide vapour, thus liquefying it, and drives it through pipe (19), round the cooling coils (7) to the float valve (11). When the liquid in the float valve reaches a certain level it causes the hollow steel float (12) to float, opening the needle valve (13), and allowing liquid sulphur dioxide to flow into the evaporator again. In this way the volume of liquid sulphur dioxide in the evaporator is controlled.

The whole unit is hermetically sealed, and requires no attention. It contains a thermostat, not shown in the diagram, which automatically switches the motor off and on when the temperature falls below or rises above the range indicated on the controls.

The expansive power of steam.

1 cu. in. of water at 100° C. becomes 1650 cu. in. of steam at 100° C. and normal atmospheric pressure. Engineers take as a rough rule that 1 cu. in. of water gives 1 cu. ft. (1728 cu. in.) of steam.

It is this enormous expansion of water when it becomes steam that makes it so suitable for use in an engine. We will calculate the work done when 1 cu. in. of water is converted into steam in a simple engine consisting of a cylinder with a movable steam-tight piston. Now 1 cu. in. of water in expanding to 1650 cu. in. of steam has to overcome a pressure of 15 lb. per sq. in. due to the atmosphere (neglecting the weight and friction of the piston). Suppose, for simplicity, that the area of cross-section of the cylinder is 1 sq. in. The piston will be forced a distance of 1650 in.

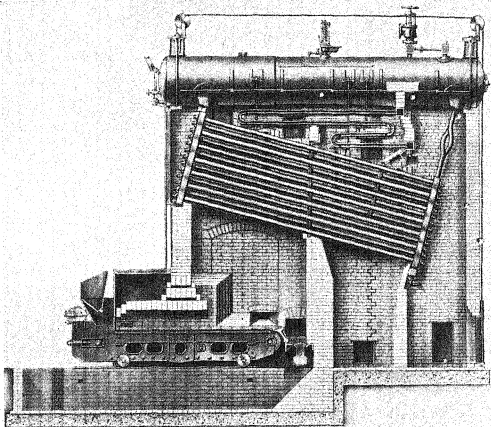
$$\text{Work done} = 15 \times 1650 \text{ in. lb.} = \frac{15 \times 1650}{12} = 2060 \text{ ft. lb.}$$

In other words a weight of nearly 1 ton could be lifted through 1 ft. This energy is derived from the source of heat.

Water-tube boiler.

In order to generate steam as efficiently as possible a boiler should be designed (i) to present a large heating surface to the flames, (ii) to cause the water to circulate rapidly and smoothly by convection (see Chapter VIII) so that fresh water is constantly receiving heat from the furnace.

The principle of the water-tube boiler is shown in Fig. 56. The water circulates through inclined steel tubes up into the steam and water drum at the top. The flames and hot gases from the furnace pass up between the tubes; they are then deflected down and up again by brickwork so that they give up as much of their heat as possible before rising up the chimney.



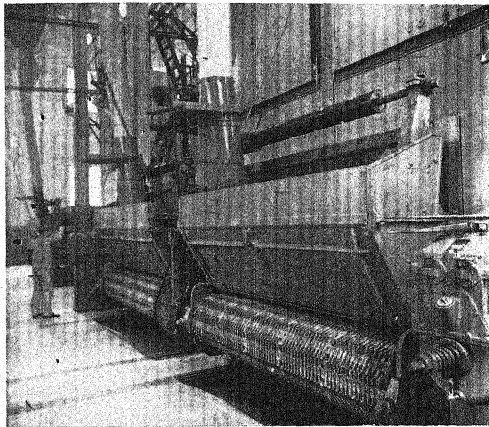
By courtesy of Messrs Babcock and Wilcox, Ltd.

Fig. 56. A drawing of the interior of a water-tube boiler, side view. The water is heated in steel tubes inclined at 15° to the horizontal and connected to a steam and water drum at the top. Coal is fed into the furnace by a moving chain grate stoker (bottom left-hand corner—see also Fig. 57). Baffles are placed between the tubes so that the furnace gases must pass up, down, up and down again before passing out into the flues (bottom right-hand corner). Steam from the drum passes through the super-heater (the single bent tube at the top), where its temperature is raised between 100° F. and 300° F. above the boiling point. Such a boiler has an evaporative capacity of 36,000 lbs. of water per hour.

SUMMARY

The latent heat of fusion of ice is the heat absorbed when unit mass of ice is converted into water at the same temperature.

Its value is 80 cal. per gm. or 144 B.Th.U. per lb.



By courtesy of Messrs Babcock and Wilcox, Ltd.

Fig. 57. The front of the water-tube boiler shown in Fig. 56. Pulverised coal is fed from the hoppers on to the moving chain grate stokers.

The latent heat of vaporisation of water is the heat required to change unit mass of water at its boiling point into steam at the same temperature.

Its value is 536 cal. per gm. or 965 B.Th.U. per lb.

When a liquid evaporates it extracts its latent heat of vaporisation from itself and its surroundings.

When a liquid is cooled its temperature remains constant during solidification owing to the liberation of latent heat. This fact may be utilised in determining the melting point.

Most substances contract on changing from the liquid to the solid state. A few substances such as water and iron, however, expand. 10 c.c. of water \rightarrow 11 c.c. of ice.

Pressure lowers the melting point of water and of all substances which expand on solidification but raises the melting point of substances which contract on solidification.

Impurities such as salt lower the melting point of ice.

A liquid continues to evaporate into the space above it until the vapour reaches a certain pressure called the **saturated vapour pressure** or **vapour pressure of the liquid**. The vapour pressure of a liquid increases with temperature.

A liquid boils when its vapour pressure is equal to the external pressure.

If two gases or vapours occupy a space, each exerts the same pressure that it would exert in that space if the other were absent. (Dalton's law of partial pressures.)

Increase of the external pressure and also dissolved impurities both raise the boiling point of a liquid.

The most common type of refrigerators utilise an easily vaporisable liquid such as ammonia, carbon dioxide or sulphur dioxide, which are made to evaporate in the low pressure part of the machine, thereby extracting their latent heat from the surroundings. They are then made to liquefy in the high pressure part, where they give out their latent heat.

1 cu. in. of water \rightarrow 1650 cu. in. of steam at 100°C . and normal atmospheric pressure.

QUESTIONS

(Take the latent heat of fusion of ice as 80 cal. per gm. and the latent heat of vaporisation of water as 540 cal. per gm. unless otherwise stated.)

1. Under what circumstances can heat be supplied to a substance without causing a rise in temperature? Can this heat be recovered?
2. Define latent heat of fusion of ice. Find the heat required to convert

(a) 10 gm. of ice at 0°C . to water at 0°C .;

(b) 15 gm. of ice at 0°C . to water at 35°C .;

(c) 20 gm. of ice at -10°C . to water at 20°C . (sp.ht. of ice = 0.5);

(d) 2 lb. of ice at 32°F . to water at 32°F .

(Latent heat of fusion of ice = $144\text{ B.Th.U. per lb.}$)

3. Define latent heat of vaporisation of water. Find the heat given out when:

(a) 5 gm. of steam at 100°C . condense to water at 100°C .;

(b) 8 gm. of steam at 100°C . condense to water at 20°C .;

(c) 1 lb. of steam at 212°F . condenses to water at 60°F .

(Latent heat of vaporisation of water = $965\text{ B.Th.U. per lb.}$)

4. What is approximately the resulting temperature if 250 gm. of water at 90°C . are added to a mixture containing 60 gm. of ice and 400 gm. of water at 0°C .?

5. If 10 gm. of steam at 100°C . are blown into 200 gm. of water at 10°C ., what will be the resulting temperature of the water?

6. Pieces of dry ice were dropped into a copper calorimeter weighing 120 gm., and containing 225 gm. of water at 21°C . After the ice had melted the temperature was found to be 6°C . How much ice was used. (Sp. ht. copper = 0.1 .) (L.)

7. Steam at 100°C . is passed into a vessel containing 240 gm. of liquid at 6°C . When the temperature reaches 27°C . the steam jet is removed and, on weighing, it is found that 5 gm. of steam have been condensed. If the specific heat of the liquid is 0.56 , what is the water equivalent of the vessel? (L.)

8. Define the terms latent heat, specific heat.

400 gm. of a liquid at 16°C . are contained in a copper calorimeter of mass 100 gm. Heat is supplied at a constant rate by a bunsen burner, and it takes 5 min. to bring the liquid to its boiling point 156°C . How long will it take to evaporate the whole of the liquid? Neglect radiation losses. (Latent heat of evaporation of liquid = 69 cal. per gm.; sp. ht. of liquid = 0.47 ; sp. ht. of copper = 0.1 .) (C.)

9. How much heat is required to convert 10 gm. of ice at -5°C . into steam at 100°C .? (Sp. ht. of ice = 0.5 .) (G. & C.)

10. What will be the result of mixing 10 lb. of water at 60°C . and 10 lb. of ice at -10°C . (Sp. ht. of ice = 0.5 .) (N.)

11. Indicate the sort of disadvantages that would afflict the cook if the latent heat of steam were 1 cal. per gm. instead of 540 cal. per gm.

12. Explain fully:

(a) Water is kept cool in hot countries by the use of porous earthenware vessels through which the water is able to percolate very slowly, keeping the outside surface always moist.

(b) A bandage soaked in methylated spirit is much cooler than one soaked in water.

(c) Vegetables stored in a cellar during the winter are protected from frost by placing tubs of water in the cellar.

13. Describe Black's ice calorimeter.

100 gm. of a metal at 100°C . are placed inside a cavity of a large block of ice and it is found that 13.8 gm. of ice are melted. What is the specific heat of the metal?

14. A book of tables gives the following information about aluminium:

Coefficient of expansion	$= 0.000023$ per $^{\circ}\text{C}$.
Specific heat	$= 0.23$.
Melting point	$= 658^{\circ}\text{C}$.
Latent heat	$= 107$ cal. per gm.

Explain the meanings of these data.

Calculate the amount of heat required to melt 100 gm. of aluminium starting from 18°C . (N.)

15. (a) The latent heat of fusion of sulphur which melts at 113°C . is 9 cal. per gm.

(b) The specific heat of solid sulphur is 0.17 cal. per gm.

Explain the meaning of the above statements.

Find the rise in temperature to the nearest degree when 35 gm. of liquid sulphur at its melting point is poured into a copper calorimeter weighing 40 gm. and containing 100 gm. of water at 14°C . (Sp. ht. of copper $= 0.1$ cal. per gm.) (N.)

16. When steam at 100°C . is passed into a mixture of ice and water, it is found that 15 gm. are condensed before all the ice is melted. How much ice was there with the water when the steam began to pass? (N.)

17. 80 gm. of ice at -10°C . are added to a vessel containing oil at 30°C . The vessel weighed 120 gm. when empty and was made of copper. The final temperature was 5°C . Calculate the mass of the oil assuming that specific heat of oil $= 0.4$; specific heat of ice $= 0.5$; specific heat of copper $= 0.1$. (O. & C.)

18. A quart bottle of milk is cooled from 75°F. to 40°F. by placing it in an ice-chest. If the milk weighs $2\frac{1}{2}\text{ lb.}$ and the bottle $1\frac{1}{2}\text{ lb.}$, how much ice is melted? (Sp. ht. of milk = 1, of glass = 0.2. Latent heat of fusion of ice = 144 B.Th.U. per lb.) (L.)

19. In an ice-making machine water is frozen by evaporating ammonia. Calculate how much ammonia must be evaporated at the freezing point of water to form 1 cwt. of ice from water at 50°F. (Latent heat of vaporisation of ammonia = 568 B.Th.U. per lb.; latent heat of fusion of ice = 144 B.Th.U. per lb.) (L.)

20. In a shot tower, hot solid lead is cooled by falling into a tank of water at 100°C. If the shot is at 150°C. and 1000 gm. fall each second, find how much water per hour will be required to replace that lost by evaporation. (Sp. ht. of lead = 0.08.) (O.)

21. Describe a simple form of ice calorimeter. What are its advantages and disadvantages?

260 gm. of sulphuric acid at 16°C. are placed in a glass vessel standing in an ice calorimeter, and 33.8 gm. of ice are melted. 100 gm. of water at 16°C. are then added, and a further 140 gm. of ice are melted. Find the specific heat of sulphuric acid, and also the heat evolved on mixing the above quantities of acid and water at 0°C. (O.)

22. A piece of iron, of mass 20 gm. and specific heat 0.089 is cooled to the temperature of liquid air and is then placed in a beaker of ice-cold water. 4.08 gm. of ice are formed round the iron. Calculate the temperature of the liquid air. (O. & C.)

23. If 4 gm. of solid Wood's metal at 15°C. are added to 72 gm. of the liquid metal at 100°C. the resultant temperature of the mixture is 85°C. and the mixture is all liquid. Assuming the specific heat of the metal to be 0.04 in both solid and liquid form, calculate its latent heat of fusion. (C.)

24. 1 gm. of ice at -10°C. is changed to steam at 110°C. Plot a rough graph of temperature and calories absorbed. (Sp. ht. of ice = 0.50; sp. ht. of steam = 0.47.)

25. Heat is applied at a constant rate to solid paraffin wax, contained in a vessel, until its temperature is a few degrees above melting point. Describe and explain, with the aid of a rough graph, how the temperature of the wax changes as the heating proceeds. (B.)

26. Explain carefully the following:

(a) When walking on snow a hard crust of snow forms under and sticks to the boot, except when the temperature is well below freezing.

(b) In an indoor skating rink the temperature of the ice is sometimes considerably below 0°C . and skating is difficult since the skate will not "bite".

(c) It is believed that the interior of the earth is solid, despite the fact that its temperature is higher than the melting point of all known substances.

(d) Salt is added to ice to form a freezing mixture but it is thrown on icy pavements in the winter to melt the ice.

27. A gram of ice at -20°C . is heated gradually until its temperature is 120°C . Describe the changes in volume and state which it would undergo. Draw a rough graph of volume against temperature. State any accurate values of volume change that you know.

28. Describe and explain experiments to prove that increase of pressure (a) lowers the melting point of ice, (b) raises the boiling point of water.

Describe two examples of, or applications of, one of these phenomena. (O. & C.)

29. A weight is suspended by a loop of copper wire passing round a block of ice. Describe and explain carefully what happens. (O. & C.)

30. What is meant by the boiling point of a liquid and on what does it depend?

If you had no barometer but had an accurate thermometer, how would you find out whether the atmospheric pressures on two different days were the same or not? If the pressure had changed, how could you estimate the extent of the change?

31. Describe fully what you would do if you wished to cool a silver thimble below 0°C .

Explain the reasons for the cooling effect in two such methods. (C.)

32. Describe two methods of finding the melting point of beeswax.

33. Fig. 58 represents Wollaston's cryophorus.

A bulb *A* containing water is connected to a bulb *B* which contains no water. *B* is surrounded by a freezing mixture and after a time the water in *A* freezes. Explain fully how this happens.

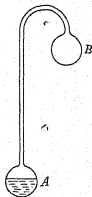


Fig. 58

34. Explain the full significance of the expression "Maximum vapour pressure of a liquid at a given temperature".

How could you measure this in the case of ether at the ordinary temperature of the laboratory? (N.)

35. A text-book of physics states that the vapour pressure (measured in cm. of mercury) of water at the temperature 50°C. is 9.20 cm. What does this mean? What experiment could you perform in the school laboratory to find the vapour pressure of water at temperatures between 15°C. and 60°C. ? (N.)

36. Distinguish between evaporation and boiling.

How would you show experimentally that (a) a liquid boils when its vapour pressure equals the external pressure, (b) dissolving a substance in a liquid lowers its freezing point. (O. & C.)

37. How would you determine the boiling point of (a) a solution, (b) a liquid of which only about 1 c.c. is available? (L.)

38. What do you understand by the vapour pressure of a liquid?

Some water was placed in a vessel of constant volume and the temperature was gradually raised while the readings were taken on a pressure gauge communicating with the vessel.

Plot a graph to show the results which were as follows:

Temperature ($^{\circ}\text{C.}$)	20	25	29	32	35	38	41
Pressure (mm.)	17.5	23.7	30.0	30.7	31.1	31.5	32.0

What information can you deduce from this graph as to the vapour pressure of water? (O.)

39. Explain the difference between a gas and a vapour.

A vessel contains dry air at a pressure of 760 mm. On admitting some water the pressure rises to 780 mm. and a small quantity of liquid water remains. If the volume of the vessel is now reduced to half its original volume, what will be the pressure? (O.)

40. Describe how you would find the vapour pressure of water at temperatures between 0°C. and 100°C.

How would you expect your results to be affected if you used in this experiment instead of pure water (a) sea water, (b) a suspension of chalk in water? (O.)

41. Explain what is meant by a saturated vapour. What are the chief properties of such vapours?

Plot a vapour-pressure curve for water from the following data:

Pressure of saturated water vapour (mm.)	720	746	773	802
Temperature ($^{\circ}\text{C.}$)	98.5	99.5	100.5	101.5

Use your curve to find the error in a thermometer which reads 99.7°C . when placed in steam on a day when the barometer stands at 767 mm. (O. & C.)

42. Define a saturated vapour. How would you show experimentally that a space may be saturated with one vapour and yet can take up the vapour of a different liquid, with an increase of total vapour pressure? A litre flask is filled with water-saturated air at a temperature of 50°C . and a pressure of 70 cm. of mercury; it is corked and cooled to 15°C .; what is the new internal pressure? (Saturated vapour pressure of water at 15°C . = 1.3 cm. of mercury; at 50°C . = 9.2 cm. of mercury.)

Would the converse of this method hold good if the change in temperature had been the other way? Why?

43. Describe and explain the changes in pressure which occur when a volatile liquid is introduced, one drop at a time, into a vacuous space. What difference would there be if the space were filled with dry air at atmospheric pressure instead of being a vacuum? (C.)

44. Distinguish between a saturated and an unsaturated vapour and illustrate your answer by explaining what happens in the following experiments:

(a) Ether is introduced a drop at a time into the space above the mercury column in a barometer tube.

(b) A long barometer tube, standing in a deep mercury trough, contains a little liquid ether above the mercury column in the tube. The tube is then pulled further out of the mercury trough. (C.)

45. Mercury thermometers are made reading to 400°C . although mercury boils at 357°C . They contain nitrogen at a pressure of 3 or 4 atmospheres above the mercury in the stem. Explain fully the physical principle involved.

Chapter VI

THE NATURE OF HEAT

In Chapter IV we described how heat is measured. We made no inquiries about its nature, but we assumed that heat can be passed from one body to another unchanged in amount. We now ask, what is heat?

The researches of Joseph Black (the first experiments in the measurement of heat) were guided by the theory that heat is an invisible, weightless fluid called *caloric*. He imagined this fluid passing from a hot body to a cold, or from a flame to a beaker of water. So much caloric was gained by one body and an equal amount lost by the other. He found this theory a useful tool, and a simple and satisfactory explanation of his experiments. Indeed the theory seemed so straightforward that it became widely accepted.

The experiments of Rumford.

At the beginning of the nineteenth century the caloric theory began to break down. Its inadequacy was exposed in a paper read before the Royal Society in 1798 by Count Rumford.

Rumford had been engaged in the boring of cannon at the military arsenal in Munich and he was struck by the large amount of heat developed. The metallic chips, the cannon and the borer all became hot, despite the fact that there was no apparent source of heat, except the friction between the borer and the cannon.

Such a phenomenon was not unknown to the supporters of the caloric theory, and they were prepared with an explanation. They said that the metallic chips had a smaller specific heat than larger pieces of metal. Accordingly, the heat originally contained in a lump of metal would cause a rise of temperature when the metal was chipped: there was no need to look for an external source of heat. It was shown by experiment, however, that there is no difference between the specific heats of brass in bulk and brass in powdered form. In face of this fact the supporters of the caloric theory had to change their ground. But they were not

defeated. They said that there was an evolution of latent heat when brass is powdered, similar to that known to occur when one form of tin changes to another.

Rumford then used a blunt borer so that no metallic chips were formed. He also surrounded the cannon with a wooden box containing $2\frac{1}{2}$ gallons of water. While the borer pressed tightly against the base of the cannon, he caused the latter to revolve by means of horses. At the end of two hours and twenty minutes the water boiled.

It seemed incredible that so large a quantity of caloric could have come out of any metallic filings that might have been formed, or even out of the cannon itself. Moreover, the supply was, apparently, inexhaustible. What then was the source of the heat? This is Rumford's own answer: "...it appears to me to be extremely difficult, if not quite impossible, to form any distinct idea of anything capable of being excited and communicated in these experiments except it be *motion*." In other words, the ultimate source of the heat was the motion of the horses. *The work they did was somehow converted into heat.*

Heat produced by compression of gases.

Anyone who has pumped up a bicycle tyre has demonstrated that work may be converted into heat: after a few strokes the barrel of the pump becomes warm. The air molecules are accelerated by collision with the piston of the pump, like balls being driven by a bat. All gases are heated in this way when suddenly compressed.

The nature of heat.

We now believe that *the heat a body contains is the motion, or to use a more exact term, the kinetic energy of its molecules*. At the temperature of the absolute zero, when a body has no heat at all, its molecules will stop moving.

In Rumford's experiments the friction between the borer and the metal caused the molecules of both borer and cannon to vibrate more quickly. The heat developed was actually the increase in the kinetic energy of the molecules. When the cannon was surrounded by water, the brass molecules of the cannon passed on some of their energy to the water molecules.

The idea that heat was a form of motion had strongly attracted the more profound of the earlier natural philosophers, men like Bacon, Newton and Boyle. Indeed, Bacon wrote plainly "Heat itself is motion and nothing else". But the caloric theory was simpler and the experimenters held on to it as long as they could.

Mayer and Joule.

Although Rumford had dealt the caloric theory a mortal blow it was a long time dying. It held its ground for another fifty years until, at length, Joule administered the *coup de grâce* by firmly establishing its rival.

One of the foremost opponents of the caloric theory and champion of the kinetic nature of heat was a German, Mayer. We are told that on one occasion his friend Jolly pointed out to him that, if his views were correct, a flask of water could be warmed by shaking—a *reductio ad absurdum*! Mayer admitted the conclusion and then proceeded to demonstrate that it was actually a fact.

Indeed the water at the foot of a waterfall is at a measurably higher temperature than that at the top. For after dropping and acquiring a considerable speed it is brought suddenly to rest. The organised motion of the molecules as a whole is converted into random motion of the molecules, i.e. heat. Joule was observed during his honeymoon in Switzerland, carrying a large thermometer with which to detect the slight difference between the temperatures of the water at the bottom and top of the high waterfalls there.

The Mechanical Equivalent of Heat.

The final overthrow of the caloric theory was accomplished by the experiments of Joule. A wealthy Manchester brewer, he had his own private laboratory, and was a distinguished example of the English tradition of "amateur" scientist—a type now almost extinct. Most modern research is carried out by a team of experimenters—such as that at the Cavendish Laboratory, Cambridge, directed by Lord Rutherford.

Joule's ideas gained prestige through the support of William Thomson, afterwards Lord Kelvin, whose distinguished career at Cambridge had brought him fame. Joule records of his paper on "The Mechanical Value of Heat", read before the British Associa-

tion in 1843: "The subject did not excite much general attention; so that when I brought it forward again at the meeting in 1847, the chairman suggested that, as the business of the Section pressed, I should not read my paper, but confine myself to a short verbal description of my experiments. This I endeavoured to do, and discussion not being invited, the communication would have passed without comment if a young man had not risen in the Section, and by his intelligent observations created a lively interest in the new theory. The young man was William Thomson."

Joule produced heat from mechanical work in many ways, e.g. (1) churning liquids such as water, oil and mercury, (2) driving water through fine tubes, (3) the friction between cast-iron bevelled wheels, and (4) passing an electric current through a coil of wire immersed in water.

But where he made such a significant advance was that he *measured* the heat developed and he *measured* also the work producing the heat. And he found that *a certain amount of work always produced the same quantity of heat.*

Often the heat developed was so small that it caused a rise of temperature in his calorimeter little more than half a degree Fahrenheit. Critics complained that Joule drew big conclusions from small observational effects. But he estimated his temperatures to one-thousandth of a degree Fahrenheit, a very unusual degree of accuracy in those days, and he repeated his experiments until his results were consistent. The experiment which we shall now describe was repeated forty times.

A brass paddle wheel with eight sets of vanes moved between four sets of fixed vanes fitted inside a large calorimeter containing water (see Fig. 59). A very considerable effort was necessary to turn the paddles, since, by reason of the fixed vanes, the water was churned and did not swirl round as a whole. The churning of the water generated heat which was measured by weighing the calorimeter and the water, and observing the rise in temperature. Joule used a thermometer reading to $\frac{1}{100}^{\circ}$ F.

The paddle was turned by known weights, W , falling through a measured distance, and the work done was measured by the product of the weights and distance. The weights were allowed to fall about twenty times and after each descent they were wound up on a drum D which was disconnected from the axle of the paddle wheel.

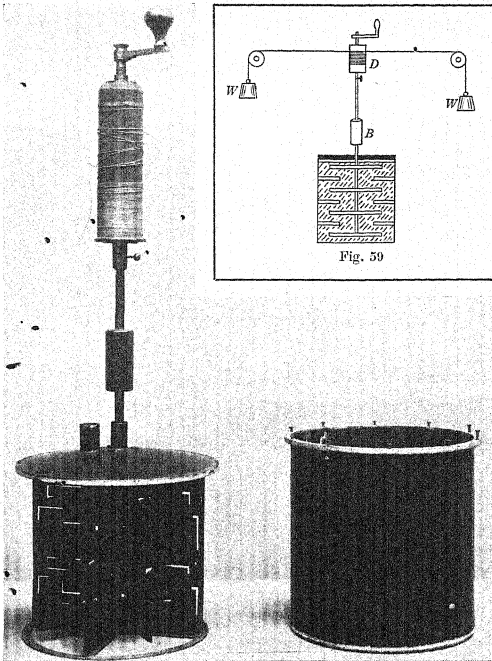
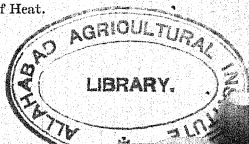


Fig. 59

By courtesy of the Science Museum.

Fig. 60. Joule's original calorimeter, employed in 1849 in his determination of the Mechanical Equivalent of Heat.



Joule took careful precaution to prevent loss of heat from the calorimeter during the experiment. Loss by conduction along the axle of the paddle wheel, for instance, was minimised by making a portion of the axle of boxwood, *B*. A correction was applied for loss of heat by cooling.

Further corrections were made for the unused kinetic energy of the weights when they bumped the ground, friction in the pulleys, the strain energy stored in the cords due to stretching and given to the paddle when the weights reached the ground, and even for the energy wasted in the humming sound generated.

Let W lb. = falling weights,

h ft. = distance fallen,

n = number of descents,

M lb. = mass of water,

m lb. = water equivalent of calorimeter and paddle,

t° F. = rise in temperature.

Then

Work done = Whn ft. lb.

Heat generated = $(M + m) t$ B.Th.U.

Now the amount of work equivalent to 1 unit of heat is known as the **Mechanical Equivalent of Heat**.

$$\therefore \text{Mech. Equiv. Heat} = \frac{Whn}{(M + m)t} \text{ ft. lb. per B.Th.U.}$$

Joule obtained as his final value of the Mechanical Equivalent of Heat, often denoted by J in his honour, 772 ft. lb. per B.Th.U.

The experiment was repeated by Prof. Rowland, an American, who introduced several improvements, and, in particular, used more power, which enabled him to obtain a greater rise of temperature than in Joule's experiment. In 1879 he published the value, now generally accepted,

$$J = 778 \text{ ft. lb. per B.Th.U.}$$

or

$$4.18 \times 10^7 \text{ ergs per cal.}$$

The Principle of the Conservation of Energy.

Towards the middle of the nineteenth century a new idea was germinating in the minds of such men as Mayer, Joule and William Thomson, and there gradually evolved one of the most fundamental generalisations of science, the Principle of the Conservation of Energy.

It was becoming clear that not only could work be converted into heat, and *vice versa*, but electricity could be converted into heat, heat into light, work into sound, and so on: in fact, all were mutually convertible.

A new word was coined, "energy", to represent that of which work, heat, light, sound, electricity, etc. were different forms.

Energy may be defined as the capacity to do work. Its real significance, however, can best be made clear by an analogy. A sum of money, representing a definite amount of wealth, may be converted into different currencies, pounds, dollars, francs, marks or roubles. The real wealth of the sum of money corresponds to energy, and the currencies to its different forms, heat, electricity, and so on.

The experiments of Joule established the fact that when work is converted into heat there is always a fixed rate of exchange, known as the Mechanical Equivalent of Heat. Hence arose the idea that energy, when it is converted from one form to another, is never lost or destroyed. This is the *Principle of the Conservation of Energy*, namely, *energy can neither be created nor destroyed*.

A vivid example of the meaning of the principle was given by Prof. Tyndall, famous for his gifts as a popular lecturer. Suppose a sledge hammer is allowed to fall on to a lump of lead. If it were possible to gather up all the heat and sound energy caused by the blow there would be exactly sufficient to raise the hammer back to its starting point.

Similarly an express train, brought to rest by the application of its brakes, loses all its energy of motion. But the heat generated at the brakes, the rails, and in the air, together with the energy of the sound emitted by the brakes, is exactly equivalent to the original kinetic energy of the train.

A common series of transformations of energy is as follows:

Chemical energy → heat → motion → motion of electricity → heat → light
 coal steam engine dynamo bulb

All of the energy of the coal is by no means converted into heat and light in the electric-light bulb. But the missing portion has not been destroyed, only wasted in the form of heat which has escaped into the atmosphere.

The principle of the Conservation of Energy is really another form of the belief, held by scientific men since the days of

Galileo, that a perpetual motion machine is impossible. No machine can create energy; hence, since it must always do work against friction, it cannot continue to run indefinitely without its supply of energy being replenished.

SUMMARY

Heat is a form of energy.

The heat a body contains is the kinetic energy of its molecules.

The mechanical equivalent of heat is the amount of work equivalent to 1 unit of heat.

Its value is

$$4.18 \times 10^7 \text{ ergs per calorie,}$$

or

$$778 \text{ ft. lb. per B.Th.U.}$$

Energy can neither be created nor destroyed (Principle of the Conservation of Energy).

The caloric theory was overthrown and the kinetic theory of heat established by the experiments of Rumford, Mayer and Joule.

QUESTIONS

1. Give an account of the experiments performed by Rumford, Davy and Joule which prove that heat is a form of energy and not a material substance. (O. & C.)

2. State the evidence which led Rumford to assert that heat was not a form of matter.

Describe any experiment which seems to you to be conclusive proof that heat is a form of energy, pointing out clearly why it is that you consider the proof to be conclusive. (O.)

3. What is the vibration theory of heat?

Describe in terms of this theory the effect of rise of temperature on the molecules of (a) gases, (b) liquids, and (c) solids. (O.)

4. What do you understand by the term 'energy'?

Mention four different forms (other than heat) in which energy can exist. What is the evidence that heat is a form of energy, and how would you demonstrate this in the laboratory? (O.)

5. Explain:

(a) Flint and steel, when struck together, produce sparks.

(b) When the valve of a motor tyre is opened the escaping air is cooled.

6. Indicate a method by which the mechanical equivalent of heat may be determined.

At a pumping station it was found that 1000 gallons of water were pumped to a height of 100 ft. for every 1.2 lb. of fuel used. Express as a decimal, the ratio between the work done and the energy of the fuel. (1 gallon of water weighs 10 lb.; 1 lb. of fuel \rightarrow 15,000 B.Th.U.; Mech. equiv. = 780 ft. lb. per B.Th.U.) (N.)

7. Calculate the difference in temperature of the water at the top and bottom of the Niagara Falls, the height of the falls being 160 ft.

($J = 778$ ft. lb. per B.Th.U.)

8. In a rough determination of the mechanical equivalent of heat a quantity of lead shot was placed in a cardboard cylinder of length 1 metre and was allowed to fall backwards and forwards from one end of the tube to the other 50 times in succession. The temperature of the shot was found to have risen 3.7° C. Calculate the mechanical equivalent of heat. (Sp. ht. of lead = 0.03, 1 cm. gm. = 981 ergs.)

(O. & C.)

9. Water is heated in a vessel by rotating vanes which are driven by a 10 kgm. weight falling through a height of 2 metres. If the vessel weighs 100 gm. and is made of metal of specific heat 0.2 and contains 160 gm. of water, how much will the water rise in temperature? (Mech. equiv. of ht. = 4.2×10^7 ergs per cal.; $g = 981$ cm. per sec.²)

(O. & C.)

10. If a lump of lead fell from a height of 100 ft. onto a stone pavement, by how much would its temperature be raised, supposing that all the heat caused by the blow were retained by the metal? (1 B.Th.U. = 780 ft. lb.; sp. ht. of lead = 0.03.) (N.)

11. A block of ice falls from the edge of the Ice Barrier into the water below. Both ice and water are at 0° C. As a result of the fall one-fiftieth of the ice is melted. What is the height of the Barrier? ($J = 4.2 \times 10^7$ ergs per cal.; $g = 1000$ cm. per sec.²; $L = 80$ cal. per gm.) (O. & C.)

12. If all the heat of a furnace could be turned into useful work, what would be the horse-power of a steam engine which consumed 240 lb. of coal per hour?

(1 lb. of coal gives 12,500 B.Th.U. each of which is equivalent to 770 ft. lb., and 1 H.P. = 33,000 ft. lb. per min.)

In practice there is much waste of energy; what are the chief causes of this? (N.)

13. The water in a tank is stirred by an engine generating $\frac{1}{4}$ H.P. If the tank contains 5 gallons of water, how long will it take for the temperature to rise 22° F., assuming that all the work performed by the engine is converted into heat? (1 gallon of water weighs 10 lb.; 1 H.P. = 550 ft. lb. per sec.)

14. Explain fully what is meant by the Mechanical Equivalent of Heat.

How long should a motor-cycle working at 4 H.P. be able to run on 1 gallon of petrol, if one-fifth of the energy of the fuel is converted into useful work? (1 gallon of petrol in burning yields 120,000 B.Th.U.; 1 H.P. = 33,000 ft. lb. per min.; mech. equiv. of ht. = 770 ft. lb. per B.Th.U.) (C.)

15. Explain why a friction brake cannot be used to find the horse-power of a large engine. Calculate the heat developed per min. in a brake by an engine developing 1000 H.P. (1 H.P. = 33,000 ft. lb. per min. $J = 778$ ft. lb. per B.Th.U.)

16. A motor-car travels 30 miles in an hour and uses exactly 1 gallon of petrol. If the heat of combustion of 1 gallon of petrol is 138,000 B.Th.U. and 80 per cent. of this heat is wasted, calculate the average horse-power exerted. (1 H.P. = 33,000 ft. lb. per min.)

17. A motor-car weighing 1 ton is travelling at 30 m.p.h. Its kinetic energy in ft. lb. is given by the expression

$$\frac{Wv^2}{64},$$

where W is its weight in lb. and v its velocity in ft. per sec.

Calculate its kinetic energy. If the car is brought to rest by the action of the brakes, find the heat developed in the brake drums.

18. A bullet of mass 30 gm. is travelling with such a velocity that it just melts completely when it strikes a fixed target. Find (a) the heat produced by the impact, (b) the velocity of the bullet just before the impact, assuming that the heat is all concentrated in the bullet. The temperature of the bullet before impact may be taken to be 27° C., its melting point 327° C., its specific heat 0.03 and its latent heat 5 cal. per gm. The mechanical equivalent of heat = 4.2×10^7 ergs per calorie. (C.)

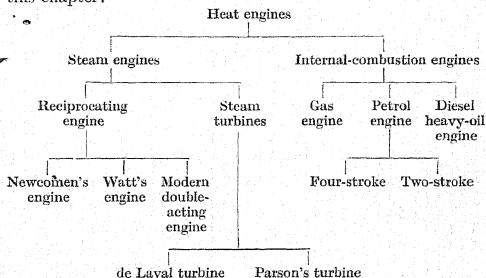
19. A thermometer has its bulb enclosed in a bicycle tyre which has been deflated. The tyre is rapidly pumped up and allowed to stand for some time. After this the valve is suddenly opened. Describe and explain the way in which the readings of the thermometer will vary. (Civil Service.)

Chapter VII

HEAT ENGINES

We have seen that mechanical work can be converted into heat. The reverse process, the conversion of heat into work is performed by heat engines.

Heat engines are conveniently grouped into two main classes: steam engines and internal-combustion engines. Steam engines again subdivide into two distinct types: reciprocating engines, in which steam causes a to-and-fro motion of a piston which is converted into rotary motion by a crank mechanism, and turbines, in which steam causes rotary motion direct. The following table shows the various types of engine which we shall discuss in this chapter:



THE STEAM ENGINE

Historical.

The steam engine, which plays so important a part in modern civilisation, was not invented by any one man. It developed gradually during the eighteenth century as a result of the genius of engineers, mainly British. Prominent among these were

Thomas Newcomen, a Dartmouth blacksmith, and James Watt, an instrument-maker at Glasgow University. Towards the end of the nineteenth century a new and revolutionary type of steam engine, the turbine, was invented by Sir Charles Parsons.

Newcomen's engine.

The first commercially successful steam engine, used for pumping water out of mines, was made by Newcomen and patented in 1705.

It consisted of a vertical cylinder and a piston, *P*, which was connected to a horizontal beam *B* (see Fig. 61). The other end of

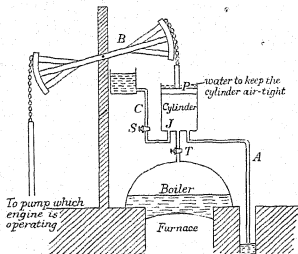


Fig. 61. Newcomen's steam engine.

the beam was connected to the piston of the pump which it operated. The weight on this side of the beam was sufficient to draw the piston *P* to the top of the cylinder. Steam was admitted from a boiler through the tap *T* to the cylinder. *T* was then closed and the steam in the cylinder made to condense by cooling the outside of the cylinder with cold water. (In Fig. 61 the cooling is done by the jet *J*, which we shall describe later.) A partial vacuum was thus caused inside the cylinder, and the pressure of the atmosphere acting on the upper side of *P* forced it down. The condensed water flowed out of the cylinder through a waste pipe *A*, which needed to have a vertical height of more than 34 ft. to counterbalance the pressure of the atmosphere at its

lower end. Steam was again admitted and the cycle of operations repeated.

Newcomen later introduced several improvements. He noticed one day that his engine made several rapid strokes in quick succession. He discovered that this was due to a hole in the piston through which the water, resting on it to make it steam tight, was percolating into the cylinder and condensing the steam much more effectively. Accordingly he arranged to condense the steam by means of a jet of cold water ejected into the interior of the cylinder through the pipe *C*, having a tap *S*.

While the engine was working someone had to operate the two taps *T* and *S* at the appropriate times. This task was entrusted by Newcomen to a boy named Humphrey Potter, who found the work monotonous. Being of an ingenious turn of mind, he tied the handles of the taps to the beam by cords in such a way that the engine itself opened the taps. The timing was now so much more perfect that the speed of the engine increased from seven or eight strokes per minute to sixteen.

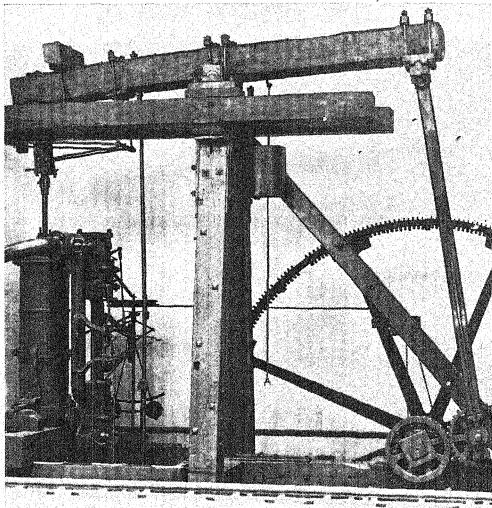
Newcomen's engines were clumsy pieces of ironmongery, and their consumption of coal was enormous. However, they served an urgent need in mining districts, and came into general use.

James Watt.

In the hands of James Watt, Newcomen's engine was completely transformed. It became, in all essential features, the steam engine we know today, and was one of the greatest contributing factors to the Industrial Revolution, the birth of the modern machine age.

Watt was born at Greenock, Scotland, in 1736. As a lad he was interested in mechanism of all kinds. In 1755, after a year's experience in London, he obtained a post as mathematical instrument maker at Glasgow University. While there, he came across a Newcomen engine which was in a state of disrepair. He set to work to improve it. He was most impressed by the large quantity of steam which it required. He approached the problem from a scientific standpoint and as a result of several experiments was convinced that there was a tremendous waste of steam due to the heat required to warm up the cylinder after each condensation. He realised that, if only the cylinder could be kept hot all the time an enormous economy could be effected. At last the

solution dawned upon him—a separate condenser. He connected the cylinder to a vessel, now known as a condenser, which he exhausted by means of a pump. On opening a tap the steam rushed out of the cylinder into the exhausted condenser where it was condensed by cold water.



By courtesy of the Science Museum

Fig. 62. Early example of James Watt's double-acting rotative engine with separate condenser and air pump. It was erected in 1788 at Boulton's Manufactory, Soho, Birmingham.

The resulting engine was much more efficient but nevertheless similar in operation to that of Newcomen, a single-acting

atmospheric engine exerting a pull only on the downstroke of the piston.

The next improvement Watt introduced was to make steam, instead of the atmosphere, force the piston down. He closed the top of the cylinder and admitted steam above the piston when the latter was at the top of its stroke. The pressure of the steam forced the piston down. The steam in the cylinder was then drawn off into the condenser by a series of valves and the weight at the far end of the beam drew the piston up again.

Watt's next invention was the double-acting steam engine, and thereafter his engines ceased to have very much resemblance to the original Newcomen engine. Steam was admitted alternately to each side of the piston with the result that the piston exerted a force during every stroke (both up and down) instead of during alternate strokes (see p. 112).

Most of the essential details of the modern steam engine were also invented by Watt. He invented the method of using steam expansively, the throttle valve for regulating the supply of steam, the centrifugal governor, the slide valve and methods of transforming the see-saw motion of the beam into rotary motion. A picture of one of his engines is shown in Fig. 62. Nowadays the reciprocating motion of the piston is converted into rotary motion by means of a crank (see p. 113) instead of a beam. Watt, however, was debarred from using this method as it was patented by one of his rivals. Watt had used the crank, but thought the idea too well known to be the subject of a patent.

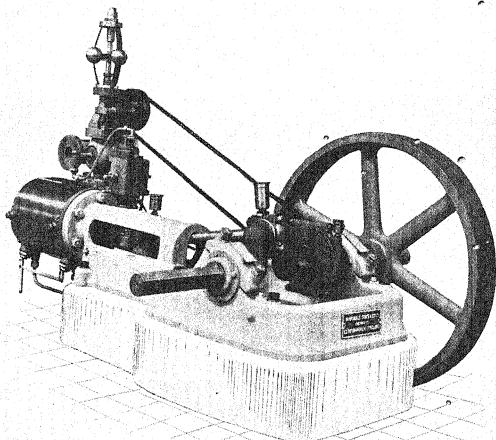
For the first forty years of his life Watt lived in very straitened circumstances. The tide turned, however, in 1773 when he entered into partnership with Matthew Boulton, the owner of an engineering works at Soho, Birmingham. At Soho, Watt's engines began to be manufactured and, as Boulton himself said, the country went "steam-engine mad".

The modern steam engine.

A simple modern horizontal engine is shown in Fig. 63. The word 'horizontal' refers to the cylinder.

The method of its operation can be followed from the diagrammatic plan, Fig. 64. The cylinder has three openings called ports, *p*, *q*, *r*, which are opened and closed by the slide valve. When the

slide valve is in the position shown, steam enters the cylinder through p and forces the piston to the right. The slide valve at the same time moves to the left and opens first the exhaust port q through which the steam leaves the cylinder, and then r



By courtesy of Messrs Marshall, Sons and Co. Ltd.

Fig. 63. A horizontal, single-cylinder steam engine. From left to right can be seen the cylinder, cross-head, connecting rod, crank and flywheel. Steam enters the aperture above the cylinder. The chain connects the crank shaft and the governor (top left): when the speed of the engine increases the balls fly out, this operates a throttle valve and cuts down the supply of steam; hence the engine is kept running at a steady rate. Connected to the crank shaft (next to the chain but hidden in this picture behind the connecting rod) is the rod which operates the slide valve.

through which fresh steam enters. The piston is now forced to the left. This is known as a double-acting steam engine since steam

is admitted to each side of the piston in turn. The steam which is generated in a boiler (not shown), enters the cylinder from the steam chest, a box fixed to the side of the cylinder.

The piston rod is connected at the cross-head, which slides backwards and forwards between guides, to the connecting rod. The latter causes the crank shaft to revolve by means of the crank. The action of the crank can most clearly be seen in Fig. 63. The crank shaft carries a heavy flywheel which drives the engine for the two short periods in each revolution when the piston can exert no turning power, i.e. when the connecting rod and crank are in a straight line (called the positions of dead centre).

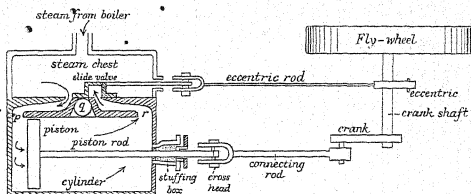
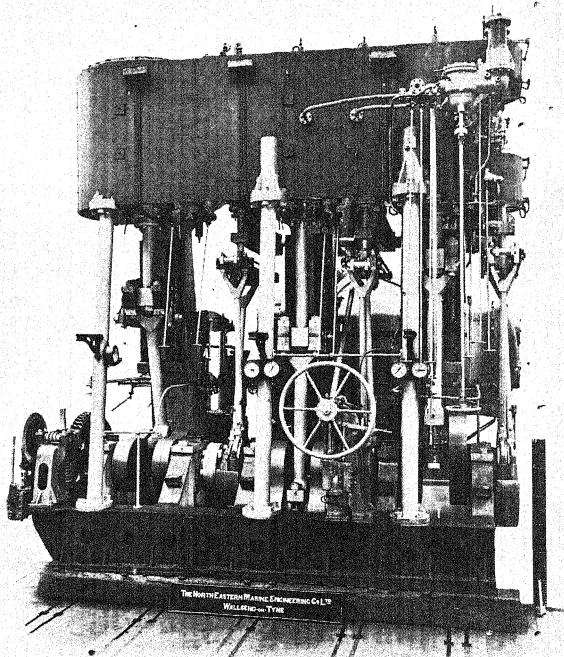


Fig. 64. Single-cylinder, double-acting steam engine.

The slide valve is connected through the eccentric rod to the eccentric, a type of crank attached to the crank shaft. It will be seen that as the piston moves to the right, the slide valve is required to move to the left.

The exhaust steam may be ejected straight into the atmosphere (non-condensing engine) or into a condenser (condensing engine). The railway locomotive is an example of a non-condensing engine. The puffs of exhaust steam are utilised for increasing the draught for the boiler. On the other hand all marine engines are condensing. There is only a limited supply of fresh water on board ship and the same water is condensed and reboiled in the engine.

The ports through which the steam enters the cylinder remain open for only a fraction of the stroke of the piston. The steam, usually superheated, is at a very high pressure, and tends to



By courtesy of the North Eastern Marine Engineering Co. Ltd.

Fig. 65. A triple expansion marine engine. The cylinders (at the top) are vertical and are supported at the front by three steel pillars. Steam enters the high pressure cylinder at the top right-hand corner of the picture and passes to the left into the intermediate and low pressure cylinders. The three

expand like an ordinary gas. Thus the expansive property of steam is made to drive the piston and an economy in steam is effected.

A simple engine has a single cylinder but a *compound* engine has two, three, or four cylinders. The same steam enters the cylinders in turn, which are of increasing bore since the pressure of steam is decreasing and its volume increasing. The steam thus suffers several expansions. The engines of a large liner are usually triple expansion engines. In this way the steam is not rejected until most of its force is spent.

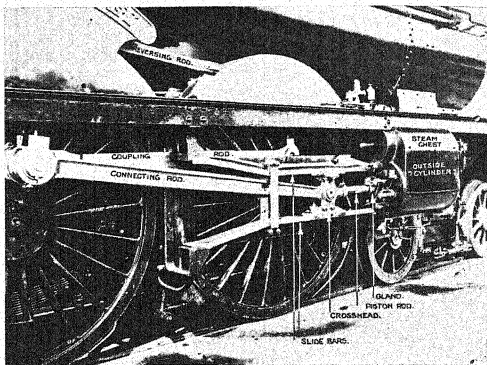
The locomotive.

The first railway was built between Stockton and Darlington in 1825 by George Stephenson. Stephenson's most famous engine was the *Rocket*, which won a competition between several engines for use on the Manchester and Liverpool railway.

The principle of the locomotive engine is very similar to that of the stationary compound steam engine. There are usually two, three, or four cylinders which are carried, two outside and the rest inside the engine frames.

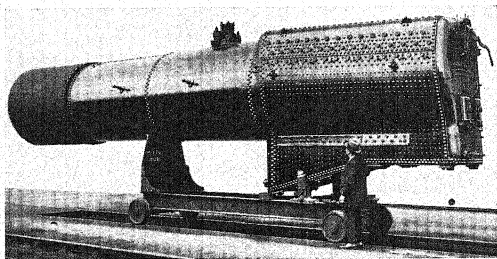
A most important and certainly the most prominent feature of a locomotive is the boiler, for the power of the engine depends on the rapid generation of steam. The type of boiler used is the fire-tube boiler. Inside the cylindrical body of the locomotive (see Fig. 67) there are a large number of small tubes through which pass the flames and smoke from the fire. The tubes are surrounded by water. In this way a very large heating surface is obtained, and steam is generated as rapidly as possible.

piston and connecting rods (nearly as thick as the steel pillars) can be seen joined at their lower ends to the crank shaft which is horizontal. The three pairs of thinner rods coupled to the crank shaft and on the right of each connecting rod, operate the valves. The left end of the crank shaft is coupled to the propeller shaft of the ship. The thin vertical rod on the right, terminating at its lower end in a horizontal wheel, is connected at the top to the throttle valve and enables the engineer to control the supply of steam. The prominent wheel at the front controls a small subsidiary steam engine, which is situated between the middle and right-hand pillars and works the reversing gear. A ship's engines must be reversible at a moment's notice. The lever connected to the left-hand cross head drives the pump which causes sea water to circulate in the condenser (situated at the back of the engine).



By courtesy of the Great Western Railway Co.

Fig. 66. An outside cylinder, connecting rod etc. of a railway locomotive.



By courtesy of the Great Western Railway Co.

Fig. 67. A locomotive boiler just ready to drop into the frames. The safety valve and clack-boxes are seen mounted on the barrel whilst the regulator handle, water gauge and fire-hole door appear on the back plate. The carrying bracket, which supports the boiler on the frames, is seen extending along the side of the firebox.

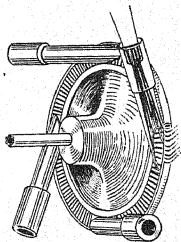
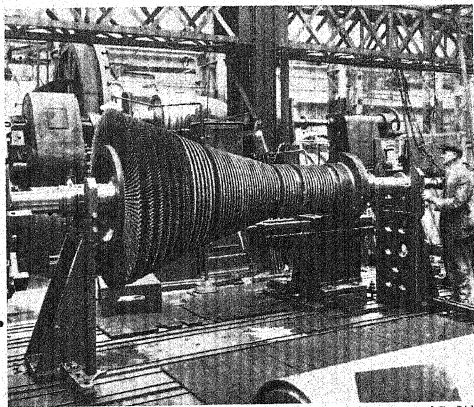


Fig. 68



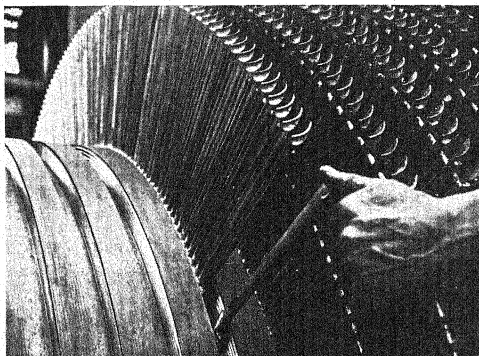
By courtesy of Messrs C. A. Parsons and Co. Ltd.

Fig. 69. The rotor of a steam turbine, completely bladed (see Fig. 70). This is for a 5000 kw., 3000 r.p.m. turbo-alternator.

The steam turbine.

So far we have considered only reciprocating engines. Nowadays, however, large liners and power plants are equipped with steam turbines similar in principle to water turbines, in which steam is made to produce rotary motion direct.

The impulse steam turbine was invented by de Laval, and in 1884 Sir Charles Parsons made the first reaction steam turbine.



By courtesy of Messrs C. A. Parsons and Co. Ltd.

Fig. 70. Method of inserting blades into the rotor of a steam turbine (see Fig. 66). The similar curved blades in the casing, which fit between these rows of blades on the rotor, face the opposite way, so that the steam is continually slewed round as it rushes through the turbine and hence turns the rotor.

Laval's turbine (see Fig. 68) is a type of steam windmill. Steam at a high pressure is directed by nozzles on to the vanes of a wheel and cause it to rotate at very high speed. The speed of rotation may be anything from 9000 to 70,000 rev. per min. and considerable gearing down is necessary.

Fig. 69 shows the rotor of a Parsons' steam turbine. It con-

sists of a number of discs or wheels of increasing size fitted with vanes and fixed to a shaft. The outer casing which is placed round this also contains fixed vanes which fit between the moving vanes. Steam at high pressure enters at the narrower end of the turbine, and passes right through it. As it impinges on the vanes it causes the rotor to revolve. The vanes are of such a shape and so arranged that the steam exerts the maximum turning force on the rotor. As the steam passes through the turbine its pressure decreases and its volume increases. Hence the vanes and the spaces between them are made progressively larger from one end to the other.

The turbine is more efficient than the reciprocating engine, and all large ships are now equipped with them. The battleship *Hood*, for instance, has turbines developing 144,000 H.P.

THE INTERNAL-COMBUSTION ENGINE

The internal-combustion engine has no fire or boiler. It derives its name from the fact that the fuel is burnt inside the cylinder itself.

The first successful internal-combustion engine was built by Etienne Lenoire in 1860. The fuel used was illuminating gas. Today stationary gas engines use coal gas, and during the war, when petrol was scarce, motor vans could be seen with a large bag ballooning on the top, carrying gas for fuel.

Liquid fuels such as petrol, alcohol, paraffin, turpentine, or even heavy oil, may be used. They require, however, to be vaporised before entering the cylinder, and this is done by a special device known as a carburettor.

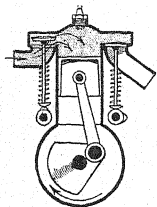
A single-cylinder four-stroke engine.

In 1878, Dr Otto of Cologne invented the four-stroke-cycle which is commonly used in gas engines and motor-car engines today.

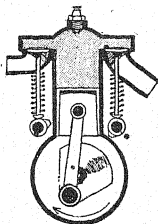
Fig. 71 shows a single cylinder. It has two valves, an inlet and an exhaust valve, of the mushroom type, which are closed by strong springs. They are opened by cams which force each up at the appropriate times.

The four-stroke cycle is so called because there is one explosion in every four strokes.

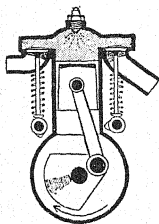
Inlet stroke. The piston moves down the cylinder; the intake valve opens and admits a mixture of gas and air.



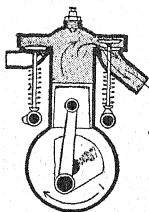
(a) Inlet stroke.



(b) Compression stroke.



(c) Explosion stroke.



(d) Exhaust stroke

Fig. 71. The cycle of operations in a four-stroke engine (single cylinder).

Compression stroke. The piston moves up and compresses the mixture. At the top of the stroke a spark passes between the points of the sparking plug and explodes the mixture.

Explosion stroke. The explosion drives the piston down; this is known as the working stroke.

Exhaust stroke. The piston moves up; the exhaust valve opens and the exhaust gases are driven out of the cylinder.

This cycle of operations is continually repeated.

The to-and-from motion of the piston is converted into rotary motion by means of a crank. On the crank shaft is a heavy fly-wheel which stores up the energy from each working stroke and drives the piston during the other three strokes. There is clearly one working stroke for every two revolutions of the flywheel.

With two cylinders there is a working stroke for each revolution, with four cylinders a working stroke for each half revolution, and so on with six or eight cylinders.

Geared to the crank shaft are two other shafts; on these are mounted the cams which strike the valve rods to open the valves. There is also a cam to time the electric spark, which is usually generated by a magneto.

The crank shaft and cam shafts are enclosed in a crank case containing oil. The connecting rod splashes the oil into the cylinders, and so lubricates the engine.

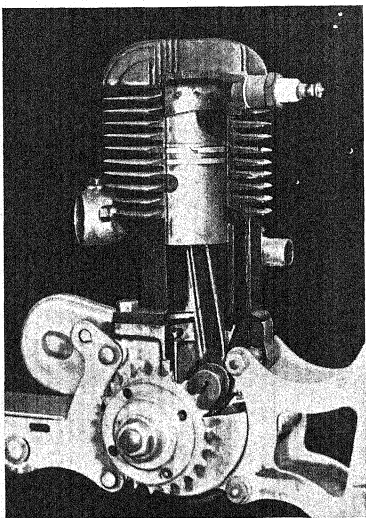
Since the fuel is burning inside the engine a considerable amount of heat is evolved, and the cylinders are cooled either by air, when they carry flanges on the outside to present as large a radiating surface as possible (see those in Fig. 72) or by water. In a stationary gas engine a continuous supply of cold water passes through the water-jacket surrounding the cylinders, but in a motor-car engine a certain quantity of water is used unchanged. It circulates through the fine tubes of the radiator where it is cooled by a strong draught of air, usually drawn through by a fan geared to the engine.

The exhaust gases pass into the atmosphere through a silencer.

Two-stroke engine.

Some motor-bicycles are fitted with two-stroke engines in which there is one explosion stroke every two strokes. Fig. 72 shows a single-cylinder two-stroke engine. The petrol vapour enters at the inlet port near the bottom of the cylinder on the right. It passes down the hollow channel called the transfer port, into the crank case which is airtight. When the piston reaches the bottom of its stroke, it uncovers the upper opening of the transfer port, and the petrol vapour in the crank case (which has been compressed) rushes up into the top of the cylinder. The

curved shape of the right-hand half of the piston is to direct the vapour up into the cylinder and away from the exhaust port—opening on the left.



By courtesy of D. G. A. Dyson, Esq.

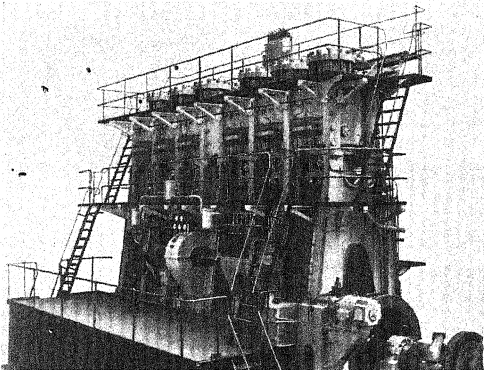
Fig. 72. Section of a single cylinder two-stroke engine. (See page 121.)

The piston now rises and compresses the petrol vapour in the cylinder. When it is at the top of its stroke the spark passes. (Note the sparking plug on the right at the top. The piston is forced down (the explosion stroke) and when near the bottom of

its stroke it uncovers the exhaust port on the left, enabling the burnt gases to escape. Meanwhile a fresh charge of petrol has been admitted into the crank case and the cycle is repeated.

The Diesel engine.

Besides gas engines burning illuminating or blast-furnace gas, and petrol engines used mainly in motor-cars and aircraft, there



By courtesy of Messrs Vickers, Ltd.

Fig. 73. A Diesel engine of 4000 brake horse power.

is a type of internal combustion engine, known as the Diesel engine which uses heavy oil as fuel. Diesel engines are used in submarines and ships, locomotives and power plants. The oil is blown into the cylinders by compressed air, where it burns during the working stroke, without the aid of a spark, owing to the high temperature of the air, taken in during the inlet stroke, consequent on its high compression during the compression stroke.

Diesel engines are more economical than petrol engines, but they are heavier. Originally invented in 1892 by Dr Rudolph Diesel, they are still in a state of development but are rapidly finding favour in many fields.

SUMMARY

See table on p. 107.

QUESTIONS

1. Describe the action of either (a) a steam engine, or (b) an oil engine.

Explain for the engine you describe how the heat of the fuel is converted into energy, and how this energy is converted into mechanical work. (O.)

2. Describe with a diagram a simple form of internal-combustion engine, and explain the principle on which it works. (O.)

3. Name three types of heat engine in modern use. What type do you consider most efficient and why?

Describe the processes involved in the transformation of heat to work in any one of the types you have named. (L.)

4. Explain carefully the function of the following parts of a motor-car engine, (a) the carburettor, (b) the valves, (c) the magneto, (d) the flywheel.

5. Superheated steam enters the cylinder of a steam engine and forces the piston to one side. Would you expect any change in temperature in the steam? Give your reasons.

Chapter VIII

THE TRANSMISSION OF HEAT

Heat is transmitted from one place to another by three separate methods, **convection**, **conduction** and **radiation**.

CONVECTION

Take a large flask two-thirds full of water and drop to the bottom one or two tiny pieces of potassium permanganate. Heat the flask with a small flame at the bottom. Coloured streams of potassium permanganate solution will be seen moving upwards and then downwards as shown in Fig. 74. These streams are called convection currents.

The convection currents are due to the fact that when a liquid is heated it expands and becomes lighter or less dense. The colder heavier liquid from above sinks and forces the warmer liquid upwards. This cold liquid in turn becomes heated and there is a continual circulation of liquid. The whole liquid therefore becomes warm because the heat is carried throughout its bulk by the moving streams or convection currents. One of the best definitions of convection is that of Clerk Maxwell: "*Convection is the motion of the hot body itself carrying its heat with it.*"

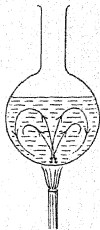


Fig. 74

Hot-water systems.

Another experiment showing convection currents may be performed with the apparatus of Fig. 75(a). When the flask is heated convection currents are set up in the direction shown by the arrows and can be rendered visible by a little colouring matter placed in the upper vessel. Can you explain why the convection currents do not circulate in the opposite direction?

This apparatus may be regarded as a model of the ordinary domestic hot-water system for heating bath water. Follow care-

fully the convection currents shown by the arrows in Fig. 75(b). The water is heated in a boiler *B*, situated as a rule behind the kitchen fire, which must be lower than the tank *C*. As soon as the fire is lit convection currents commence to circulate. The tank *A* provides the pressure which forces the water out of the bathroom taps and also replaces any water drawn off. *D* is a safety pipe in case the water in the system boils.

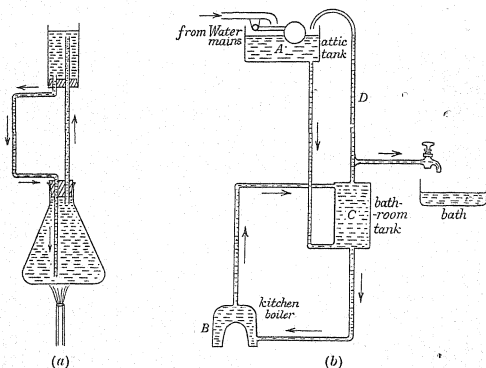


Fig. 75

Draw a diagram of a central heating system, putting in a boiler in the basement, a radiator on each of two floors, a cold-water supply cistern and an overflow pipe. Show, by means of arrows, the direction in which the water circulates.

Convection currents in the air.

The apparatus shown in Fig. 76 may be used to demonstrate convection currents in the air. It consists of a wooden box with a glass front, having two vertical glass tubes *A* and *B* projecting from its top. A candle is lit below the bottom of *A*. Then, on

holding a smouldering taper above *B*, the smoke will be drawn down in the direction of the arrows, and will ascend *A*. The air heated by the candle rises in *A* and cold air descends *B* to take its place.

Mines free from the danger of firedamp are sometimes ventilated in this way. A large fire is kept burning at the foot of, or part of the way up, an extra shaft, with the result that fresh cold air is drawn down the main shaft.

The ventilation of houses and public buildings depends on convection. In an ordinary room the fire sets up a vigorous convection current up the chimney; lamps and people also act as minor sources of air currents. The cold air usually enters through windows or crevices, no special inlets being provided.

In public buildings, however, special inlets and outlets need to be provided and they should be so designed that the air is replaced with sufficient rapidity without causing a draught. Sometimes inlets are provided in an external wall behind a radiator so that the cold air becomes warmed as it enters the room. Another common form of inlet is a Tobin tube which delivers fresh air into the room above the heads of the occupants at a height of about 10 or 12 ft. Warm "used" air rises and outlets are provided for it in or near the roof.

Factory chimneys.

Factory chimneys are not built to great heights merely to protect citizens from the smoke and the fumes. The higher the chimney the stronger the draught, and hence the fiercer the heat of the furnace. For the pressure of the draught is equal to the difference in pressures of the column of warm air inside the chimney and that of a column of cold and denser air of equal height outside the chimney (see Fig. 77). Clearly, the greater the height of the chimney, the greater is this difference in pressure.

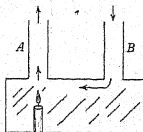


Fig. 76

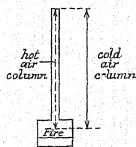


Fig. 77.

Winds.

The winds are examples of convection currents on a large scale in nature. The trade winds for instance are caused by the hot air at the equator rising and the colder air flowing in from the north and south to take its place. Owing to the rotation of the earth these winds blow from the N.E. and the S.E. It was the S.E. trade winds that blew the ships of Columbus across the Atlantic when he discovered the New World.

Land- and sea-breezes.

Near the sea coast, the wind often blows from the sea during the daytime, and from the land at night. This is primarily the result of the fact that the sea has a much higher specific heat than the land (one of the consequences of which we have already mentioned on p. 57).

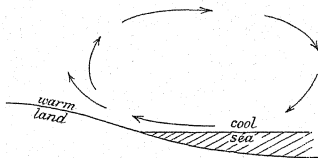


Fig. 78

The temperature of the land during the daytime rises considerably above that of the sea. The air above it becomes heated, rises, and colder air flows in from the sea, forming a sea-breeze. At night the land cools much more quickly than the sea. The air above the sea is now warmer than that above the land. It rises and cold air flows from the land, forming a land-breeze.

Other factors, besides the high specific heat of water, contribute to this phenomenon. Since the earth is a bad conductor, heat does not penetrate very far into it. Indeed the temperature a few feet below the surface remains unchanged throughout the year. On the other hand, owing to its transparency and motion, the sea becomes warmed to a considerable depth. Hence its temperature both rises and falls slowly.

Fig. 78 represents a sea-breeze blowing during the day. Draw a similar diagram showing why a land-breeze blows at night.

CONDUCTION

Coat a metal bar with paraffin wax and stick into the wax, while still warm and liquid, a number of small steel balls (see Fig. 79). The balls will stick to the bar when the wax solidifies, and will only drop off if the wax is again melted. Heat the bar at one end with a bunsen flame. The balls drop off one by one. The heat is passing from the hot end along the bar and, as the temperature of each part of the bar reaches the melting point of wax, the balls drop off.

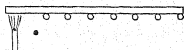


Fig. 79

This method of the transmission of heat is called *conduction*. It may be defined as "*the flow of heat through an unequally heated body from places of higher to places of lower temperature*" (Maxwell).

Different substances transmit heat at different rates, and their ability to transmit heat is known as their conductivity.

Comparison of conductivities.

A simple method of comparing the conductivities of different metals in the form of bars (of the same dimensions) is to coat the bars with paraffin wax and insert one end of each into a tank containing water (see Fig. 80). The water in the tank is boiled and the distance the wax melts along each bar enables us to classify the bars in the order of their conducting powers.

As the heat passes from the hot water through any bar, each section of the bar hands on heat to the next section and to its surroundings. In a good conductor the heat is handed on more quickly to the next section than in a bad conductor: there is therefore less time for heat to be lost to its surroundings and the good conductor is hotter all the way along.

Small steel balls may be embedded in the wax on each bar, as in the experiment described earlier, and the rate at which they drop from each bar observed. If bars of iron and lead are used it will be found that the wax melts more rapidly at first along the lead, but that

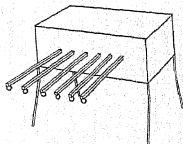


Fig. 80

ultimately it melts further along the iron bar. Now iron has a greater conductivity than lead but it has also a higher specific heat. Thus, although heat travels more quickly along the iron bar, the temperature of this bar will at first rise more slowly than that of the lead since more heat is required to raise its temperature. The wax, of course, melts as soon as the temperature of the bar reaches its melting point.

Thus, in comparing conductivities by this method we must observe, not the rate at which the wax melts, but the distance along the bar it is melted when a final equilibrium is attained.

Table of Relative Conductivities.

Silver	100	Asbestos	0.06
Copper	90	Wood	0.05
Aluminium	48	Turpentine	0.03
Brass	27	Water	0.013
Zinc	26	Cork	0.01
Iron	12	Felt	0.009
Lead	8.5	Cotton-wool	0.005
Mercury	1.5	Air	0.005
Glass	0.15		

The above table shows the relative conductivities of different substances. The conductivity of silver is taken, for convenience, as 100 and the other values are comparative.

It will be seen that the metals are far and away the best conductors of heat, while non-metals such as glass, wood and cork are very poor conductors. Liquids also are poor conductors, and gases such as air have a very low conductivity indeed.

On a cold day a good conductor always feels colder to the touch than a poor conductor, although both may be at the same temperature. This is because the good conductor conducts heat away more rapidly from the body.

A tiled bathroom floor is colder to the feet than a cork bath-mat, because tiles are a better conductor than cork. In Arctic and Antarctic regions metal implements must be covered with flannel before they are handled; otherwise, owing to their high conductivity and low temperature, they are liable to freeze to the hand.

Use of good conductors.

1. *Boilers.* Good conductors such as copper, aluminium and iron are used for cooking utensils, for hot plates on stoves, and for the construction of boilers.

The efficiency of boilers, however, is considerably reduced by two bad conductors—a layer of gas which forms on the outer surface of the boiler, and a layer of scale which is deposited on the inner surface.

Make a rectangular bag of strong paper, hang it up by threads as in Fig. 81, and pour in water. The water may be boiled by a bunsen flame if care is taken not to heat the paper above the water-line.

Clearly the paper cannot be at the temperature of the flame since it does not burn. Similarly the bottom of a boiler is not at the temperature of the furnace. There is a layer of gas (approximately $\frac{1}{40}$ in. thick), which is comparatively at rest, in contact with the paper bag or boiler, and this, having a very low conductivity can support a large difference of temperature between its opposite confines.

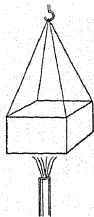


Fig. 81

The effect is most undesirable in the case of boilers since it reduces the rate at which heat is taken from the furnace. In the water-tube boiler described on p. 88 the gas film is to a certain extent removed by the "scrubbing" action of the hot gases from the furnace as they rush between the water tubes.

The second cause of inefficiency in vessels in which water is boiled, whether they be domestic kettles or large-scale boilers, is the deposit of lime and magnesia salts from hard water. It has been estimated that a layer of scale $\frac{1}{8}$ in. thick in a boiler results in an increase of three or four times the consumption of fuel, owing to its low conductivity and consequently slower rate at which heat is taken from the furnace.

The swift circulation of the water in the tubes of the water-tube boiler prevents the formation of scale in the tubes themselves. Any deposit takes place in what is called the mud-drum, from which it can easily be removed.

2. *The Davy safety lamp.* The high conductivity of copper is applied in the Davy safety lamp for use in mines.

At the beginning of the nineteenth century, when the surface workings of mines were becoming spent and shafts were sunk deeper, a series of disastrous explosions took place. It was found that the explosions were due to a gas called firedamp or methane.

Sir Humphry Davy, then Director of the Royal Institution, was consulted, and he began a series of investigations to devise a safe means of illumination for such "fiery" mines.



From Gas and Flame, the Safety in Mines Research Board Booklet. By courtesy of the Controller of H.M.S.O.

Fig. 82. A miner holding a safety lamp up to the coal face to test for gas. The flame burns blue if gas is present in a dangerous quantity. In Davy's original lamps the flame was completely surrounded by gauze and the light was feeble. In the modern form, the lower part of the gauze round the flame is replaced by thick glass, as seen in the picture, and there is a double layer of gauze at the top (within the outer metal shield), through which the air passes to the flame.

He found that methane and air explode only when raised to a high temperature. In the lamp that he invented, therefore, the flame was surrounded by copper gauze. The gauze conducted

away the heat so rapidly that the methane outside it could not reach its ignition point.

An experiment can be performed with a piece of copper gauze and a bunsen flame to show the principle of the lamp.

In Fig. 83*b* the gas has been ignited only above the gauze and will not burn below it. This is because the gauze conducts the heat away too rapidly for the gas below to reach its ignition point. Similarly in Fig. 83*a* the gas is burning only below the gauze.

Today miners are often provided with electric lamps but the Davy lamp has the advantage that when the percentage of firedamp in the mine reaches a dangerous figure, the firedamp burns blue inside the lamp and gives a warning to the miners.

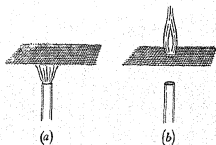


Fig. 83

Unhappily, explosions still occur in coal mines. These are often due to fine coal dust which ignites spontaneously and is therefore far more dangerous than firedamp. Another danger is carbon monoxide (CO), a poisonous gas which does not burn visibly inside the Davy lamp. In order to detect it, quick breathing creatures such as mice and canaries, which feel its effect more quickly than human beings, are taken down the mine.

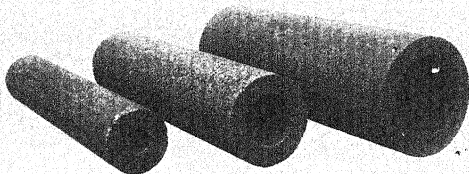
Use of bad conductors.

Heat insulation. Our clothes and the coats of animals owe most of their non-conducting properties to the air they contain. Fur and wool are the most efficient non-conductors, while hair and feathers are nearly as good. A layer of snow will often protect crops from severe frost, for snow is a poor conductor of heat, owing to the air which is trapped between the crystals.

The manufacture of materials for heat insulation is a large and important industry. Cork is commonly used for insulating cold storage chambers and refrigerating plants (see Fig. 85). It owes its low conductivity to the air which is contained in its cells, each of which is so small that loss of heat by convection is negligible.

A composition which is used for high-temperature insulation consists of 85 per cent. magnesia and 15 per cent. asbestos. The magnesia is in the form of millions of tiny crystals with air spaces

between them, giving it a very low conductivity. The asbestos, which in this case acts as a binding material, is a complex mineral consisting mainly of magnesium silicate. It has been estimated that the heat loss from 100 ft. of 6 in. bore standard steam pipe carrying steam at 100°C . represents a loss of over 200 tons of coal per annum if the pipe is left uninsulated. A layer, $2\frac{1}{2}$ in. thick, of magnesia insulation reduces this figure to about 10 tons per annum.



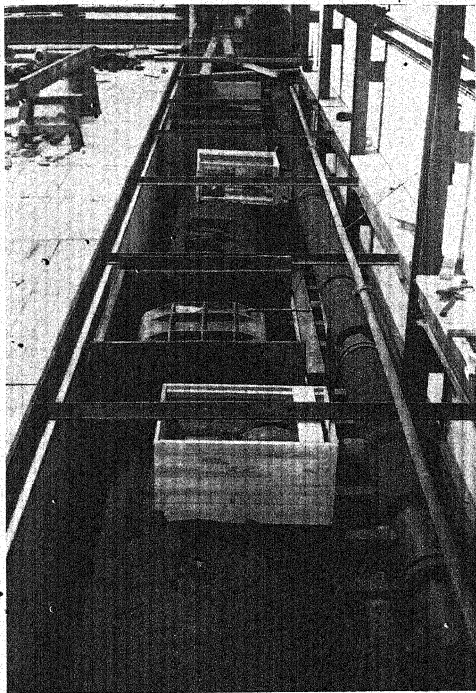
By courtesy of Messrs Mewalls Insulation Company

Fig. 84. Cork insulation used for covering pipes.

The conductivity of liquids.

Owing to the difficulty of preventing convection currents, it is much more difficult to find the conductivity of liquids than of solids. The obvious way out of the difficulty is to heat them at the top. John Dalton, in an early and somewhat crude experiment, took an ale glass full of water and inserted thermometer bulbs in the middle and at the bottom. He heated the surface with a red-hot poker, and noted that there was a slow rise in temperature recorded by the thermometers.

It may be shown that water is a very poor conductor by the following experiment. Sink a piece of ice by means of a small weight in a test-tube of water. Heat the test-tube near the top. The water may be made to boil near the surface while the ice remains unmelted.



By courtesy of Messrs Mervalls Insulation Company

Fig. 85. A pipe for carrying brine from a refrigerating plant to a cold storage chamber, in process of insulation with the cork shown in Fig. 84.

RADIATION

Heat reaches us from the sun. Since its journey of 93,000,000 miles is mainly through a vacuum there can be no question here of convection or conduction. The phenomenon is known as radiation.

Now we have seen that the heat contained by a body is actually the kinetic energy of the molecules. The energy radiated by a body, therefore, cannot be in the form of heat. A new term, *radiant energy*, is used to describe it. When it falls upon another body, it sets the molecules of the latter into more rapid vibration, and is thereby reconverted into heat.

Two theories have been put forward to explain the nature of radiant energy: (1) the Wave Theory, (2) the Quantum or Corpuscular Theory. On the former, radiant energy consists of waves in an all-pervading intangible medium called the aether; on the latter, radiant energy consists of little pellets or corpuscles of energy called quanta. We shall consider these theories in greater detail in our study of light. Light is a form of radiant energy since, on striking any body but the eye, it produces heat.

Maxwell defined radiation as follows: "*In radiation, the hotter body loses heat, and the colder body receives heat by means of a process occurring in some intervening medium which does not itself thereby become hot.*"

When a body is heated the quality of the radiation which it emits depends on its temperature. Below 500°C . it emits a form of radiation which can be detected only by its heating effect. This radiation does not, like light, excite the retina of the eye, and is therefore invisible; it is called *infra-red radiation*. (The popular term, radiant heat, is not so good, because all other forms of radiation, including light and wireless waves, have an exactly similar heating effect, and are therefore forms of radiant heat. Moreover, the radiation is not itself heat: it is only converted into heat on striking matter.)

At 500°C . a body becomes red-hot and begins to radiate red light in addition to infra-red radiation. At about 1000°C . it becomes white-hot and radiates both white light and infra-red radiation.

The difference between light and infra-red radiation, according to the Wave Theory, is that the latter has the longer wave-length;

according to the Quantum Theory, the pellets or quanta of infra-red radiation contain less energy than those of light.

All forms of radiant energy travel in straight lines at 186,000 miles per sec.; their intensity falls off with distance from the source according to an inverse square law; and they can be refracted (see the author's *Light*).

The following experiment demonstrates that infra-red radiation obeys the same laws of reflection as light.

Place a lighted electric lamp in such a position in front of a large parabolic reflector that a parallel beam of light is reflected out parallel to the axis. (The lamp is then at the principal focus of the reflector, and the arrangement is similar to the headlight of a motor-car.) Bring the light to a focus by another distant parabolic reflector (see Fig. 86).

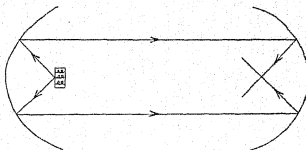


Fig. 86

Now replace the lamp by a small electric radiator. If a match-head is placed at the point where the light was brought to a focus in front of the second reflector it will be ignited, showing that the infra-red radiation has also been focussed on this point. The experiment is more convincing if an infra-red filter, which cuts out nearly all the visible radiations, is interposed between the heater and the first reflector.

Selective absorption.

Certain substances like glass absorb infra-red radiation but allow light to pass through themselves. Others like pitch permit infra-red radiation to pass, but absorb light. The phenomenon is known as selective absorption.

Glass may be used as a screen from the heat of a fire although the fire may be seen through it. This is because the radiant

energy emitted by a fire is mainly in the form of infra-red radiation.

The radiant energy of the sun, however (because of its exceedingly high temperature), is largely in the form of light. Consequently glass does not act as a screen from the heat of the sun. When sunlight passes through the glass of a greenhouse, it is absorbed by the plants and flowerpots, which become heated, and therefore begin to radiate. Since their temperatures are low, however, they radiate not light but infra-red radiation. This infra-red radiation cannot penetrate the glass and so is trapped inside the greenhouse.

Fourier suggested (about 1800) that the earth's atmosphere acts as a heat trap, thereby accounting for its equable climate. Sunlight passes through the atmosphere with negligible absorption and the infra-red radiation emitted by the earth when the latter becomes warm is absorbed by the atmosphere. On the moon, which has no atmosphere, the temperature rises during the day above that of boiling water, and falls at night below that of liquid air.

In 1859 Tyndall tested by experiment the transparency of water vapour and carbon dioxide to the infra-red radiation emitted by a cube of boiling water: he found that they were seventy and ninety times more opaque than dry air. Water vapour and carbon dioxide, therefore, are the constituents of the atmosphere responsible for the absorption of infra-red radiation.

At high altitudes there is very little water vapour and carbon dioxide in the atmosphere. The air temperature is therefore low, although sunbathing in the direct rays of sunlight may be feasible and delightful.

Arrhenius in his book, *Worlds in the Making*, suggested that in a century or two the carbon dioxide belched into the atmosphere from chimneys will render the earth's climate appreciably more equable.

The rate at which a body radiates energy.

Two factors determine the rate at which a body radiates energy: (a) its temperature, (b) the nature of its surface.

(a) Effect of temperature.

Newton discovered the following law known as *Newton's Law*

of Cooling. The rate at which a body loses heat is proportional to the difference in temperature between the body and its surroundings.

Thus a body at a temperature of 35°C. in a room with a temperature of 15°C. loses heat twice as rapidly as a body at a temperature of 25°C. (the excess temperatures being 20°C. and 10°C. respectively).

The loss of heat, however, takes place as a result of convection as well as radiation. The law is only a rough empirical rule which is not accurate for differences of temperature much greater than 30°C. between the body and its surroundings.

A more accurate law is that due to Stéfán: A body radiates energy at a rate proportional to the fourth power of its absolute temperature (T^4). This law has been the subject of considerable theoretical investigations.

(b) *Effect of the surface.*

A black body emits radiant energy more rapidly than a polished body at the same temperature.

The radiating powers of different surfaces may be compared by means of a cubical tin box into which hot water is poured. One

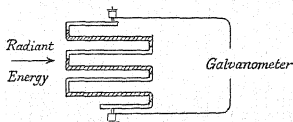


Fig. 87

of the vertical faces of the cube is lamp-blackened by holding it in the smoky flame of a candle, another polished, another painted white, and the fourth roughened. Such a box is known as a Leslie cube, after its inventor.

In order to detect and compare the radiant energy emitted by the faces of the cube, a sensitive instrument, known as a thermopile, the principle of which was described on p. 10 is used. The radiant energy is made to fall on a number of junctions of two dissimilar metals, such as antimony and bismuth, connected in series (see Fig. 87), when a small electric current is generated which can be measured by a mirror galvanometer

The junctions of the thermopile are placed at the same convenient distance from the faces of the Leslie cube in turn (see Fig. 88), and the deflections of the galvanometer noted. It is found that the order of the radiating powers of the surface is as follows: (1) lamp-black, (2) roughened, (3) white, (4) polished.

Thus, in order to minimise the loss of heat by radiation, a body should be brightly polished. A silver teapot is a good example. Hot-water radiators are painted with a dull paint to increase their radiating power, although in their case it is more important that they should have a large surface since they give out most of their heat in the form of convection currents in the air.

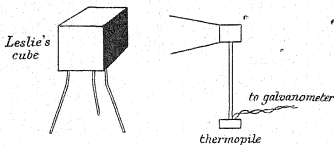


Fig. 88

The above experiment may be modified to compare the transparency to infra-red radiations of different substances.

Use the black surface of the cube and place a sheet of glass in front of the thermopile. The deflection of the galvanometer will be reduced to zero, showing that the glass is opaque to infra-red radiation. In a similar manner it can be shown that rock-salt, and also a solution of iodine in carbon bisulphide which is opaque to light, are highly transparent to infra-red radiation.

The rate at which a body absorbs radiant energy.

A black body not only radiates but also absorbs radiant energy better than a polished body. This can be demonstrated by means of Leslie's differential air thermometer (see Fig. 89), which consists of two exactly similar glass bulbs connected by a U-tube containing a coloured liquid, and also a straight tube with a tap.

Coat one of the bulbs with lamp-black and place a small electric radiator (which has not been switched on) midway between them. Open the tap *T* to equalise the air pressure in the

bulbs. The liquid will level itself in the arms of the U-tube. Close the tap *T* and switch on the heater. The level of the liquid in the right-hand arm of the U-tube will descend, showing that the pressure in the black bulb is greater than that in the polished. Thus the black bulb must be absorbing radiant energy at the greater rate.

The junctions of a thermopile are always blackened in order that they may absorb as much radiant energy as possible.

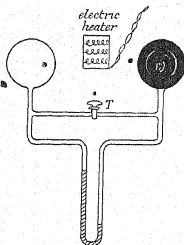


Fig. 89



Fig. 90

The thermos flask.

A vessel for storing hot or very cold liquids, e.g. hot tea or liquid air, was invented by Lord Dewar and is known as the thermos flask.

It consists of a glass vessel (see Fig. 90) with double walls, the space between which is evacuated to prevent loss of heat (in the case of hot liquids) by conduction or convection.

To minimise loss due to radiation the outside of the inner wall and inside of the outer wall are silvered: thus both silvered surfaces face towards the vacuum. The silvered surface of the inner wall radiates very little energy (a polished surface is a bad radiator) and the silvered surface of the outer wall absorbs very little of the radiant energy falling upon it but reflects most of it back. Hence there is very little loss by radiation.

EXPERIMENT. *To compare the rates at which blackened and polished surfaces absorb radiation.*

Two tin plates, about 6 in. \times 4 in., with metal pockets to hold a thermometer (see Fig. 91), are required for this experiment.

Coat the face of one of the plates with lamp-black by holding it above a candle flame, and polish the face of the other plate. Set them up vertically, with their faces parallel, at a distance of about 12 cm. apart, and place exactly midway between them an unlighted bunsen burner. Put thermometers in the pockets and read them. Note the time and light the bunsen burner (the hot blue flame should be used).

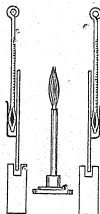


Fig. 91

Read both thermometers every minute for about 10 min., and record your readings in a table. Plot two graphs on the same graph paper, with time along the horizontal axis, and temperature upwards.

What do you deduce from this experiment? Why do the curves flatten out on top?

SUMMARY

Heat is transmitted from one place to another by convection, conduction and radiation.

1. "Convection is the motion of the hot body itself carrying its heat with it."

2. "Conduction is the flow of heat through an unequally heated body from places of higher to places of lower temperature."

3. "In radiation, the hotter body loses heat, and the colder body receives heat by means of a process occurring in some intervening medium which does not itself thereby become hot."

(Maxwell's definitions.)

Hot-water systems, winds and the action of chimneys are examples of convection.

The use of good conductors is exemplified in boilers and the Davy lamp: bad conductors such as cork and a mixture of magnesia and asbestos are used for heat insulation.

Most solids, especially the metals, are good conductors. Most liquids are poor conductors, and gases are bad conductors.

Glass, water vapour and carbon dioxide allow light to pass through them, but absorb infra-red radiation.

Newton's Law of Cooling.

The rate at which a body loses heat is proportional to the difference in temperature between the body and its surroundings.

Black bodies both absorb and radiate radiant energy more readily than polished bodies.

QUESTIONS

1. Distinguish briefly between conduction, convection, and radiation of heat, giving an example of each. Explain the part played by each in heating a building by an ordinary hot-water system. (N.)

2. Explain:

(a) In calm weather, a clear night is usually cooler than a cloudy one.

(b) The polished fire-irons in front of a fire are sometimes found to be only slightly warm, while the blackened fender is too hot to be touched.

(c) In winter you place blankets on a bed; in summer ice is wrapped in blankets.

(d) The effects of night frosts in fruit-growing districts are sometimes reduced by covering the orchards with a thick layer of smoke.

(e) There is often a draught in church pews which are below windows, even when the windows are shut.

(f) On going to bed at night we are not suffocated by the air we exhale.

(g) One end of a rod is made of copper and the other of wood. Paper is wrapped round it, and licked over with a bunsen flame. Part of the paper chars and part does not.

3. How would you compare the heat conductivities of rods of different metals? Suggest a modification of the experimental arrangements to compare the heat conductivities of liquids. What sources of error would have to be guarded against? (O.)

4. Describe and explain the action of (a) a practical appliance where use is made of the good conducting power of a material, (b) an appliance where a bad thermal conductor is employed. (L.)

5. A small garage can be heated by a Davy safety lamp. Describe this lamp and point out how its action and use illustrate methods of transference of heat. (L.)

6. Explain: (a) When rods of iron and aluminium of the same dimensions are coated with wax and heated at one end, the wax melts initially at a quicker rate along the iron rod than along the aluminium rod, but ultimately more wax is melted along the aluminium than along the iron.

(b) The temperature at the top of a mountain is less than at the foot.

(c) A gas-filled electric lamp feels hotter than an evacuated lamp when lit. Why is the filament of the former packed into as small a space as possible while the latter zigzags throughout the length of the bulb?

7. Two long rods of the same diameter, one of iron and the other of lead, are coated with paraffin wax and have small metal spheres attached to their lower surfaces by the wax.

Two ends, one of each rod, are then heated together and the spheres drop off as the wax melts. Describe and explain the order in which the spheres drop off.

	Iron	Lead
Thermal conductivity	0.14	0.08
Specific heat	0.11	0.03
Specific gravity	7.5	11.4

Sketch the arrangement of apparatus you would fit up to carry out this experiment. (L.)

8. Distinguish clearly between transmission of heat by conduction and convection. Describe, giving illustrative diagrams, how these processes are applied in a water-tube boiler. (L.)

9. What is Newton's Law of Cooling?

A vessel containing a hot liquid is placed inside a larger vessel and allowed to cool. What factors determine (a) the rate of loss of heat, (b) the rate of fall of temperature? (O. & C.)

10. Explain carefully what is meant by saying that the conductivity of copper for heat is twice that of aluminium.

Describe an experiment whereby the conductivity of a metal for heat may be found. What are the principal sources of inaccuracy in the method which you describe? (O.)

11. Explain how convection currents are set up in liquids and gases. What part does convection play in the formation of (a) land and sea-breezes, (b) trade winds?

Describe experiments to show that heat may readily be transferred by convection through water, but only with difficulty by conduction.

12. Describe an experiment by which you would show that the conductivity for heat of a specimen of paraffin oil is very small. Explain carefully the precautions you would use to ensure that the little of the heat transference actually observed in your experiment was due to convection.

13. Describe an experiment by which one could show that copper is a better conductor for heat than iron. (O.)

A small oven is being heated by a gas burner placed beneath it, and it is found that the temperature of the inside of the oven is more nearly the same all over when its walls are made of alternate layers of a good conductor such as copper and a bad conductor such as asbestos, than when the walls are made entirely of copper; explain the reason for this.

14. Describe the part played by convection in (a) some form of domestic heating, (b) the freezing of a pond. (O.)

15. What is the advantage of a high factory chimney? Why does a domestic fire often "smoke" when the fire is being lit and the chimney is cold? (O. & C.)

16. How may a cold storage cupboard be insulated? Should the ice be placed at the top or bottom of the cupboard? Explain fully.

17. Describe the various ways in which an open calorimeter containing hot water cools to the temperature of its surroundings. Explain the means you would adopt to retard the rate of cooling by each of the processes you describe.

18. Distinguish between radiation and convection of heat. Describe the conditions which give rise to (a) increased radiation, (b) increased convection. (L.)

19. It is stated that an oil stove heats a room mainly by convection, and an electric fire mainly by radiation. What experiments or observations would you make to test this statement? (L.)

20. Describe experiments to show that radiant heat can be reflected, and that it obeys the same laws of reflection as light. (C.)

21. How would you compare experimentally the radiating powers of different surfaces? Describe the Dewar ("thermos") flask and briefly explain the principles on which its action depends. (C.) (N.)

Chapter IX

THE WATER VAPOUR IN THE ATMOSPHERE

We have seen in Chapter V that the possible amount of water vapour in the atmosphere is limited by its saturated vapour pressure and that this amount increases as the temperature rises. If the water vapour is saturated and its temperature is lowered, some of it condenses in the form of drops. We shall see how this process gives rise to the various features of the weather, dew, fog and mist, clouds, rain, snow and hail.

Dew.

Early on an autumn morning a layer of dew can often be seen to cover the grass. Dew is most copiously deposited in the autumn because at that season of the year cold nights follow warm days. The water vapour in the atmosphere during the daytime is more than sufficient to cause saturation when it is cooled at night; during the night, therefore, the water vapour condenses into drops. Dew is formed on the ground because the ground loses heat by radiation and then cools the air above it.

The temperature to which air must be cooled for dew to be deposited is known as the dew-point. At this temperature the water vapour present in the air is saturated.

The atmosphere can always be made to yield dew by artificially cooling it, and the experiment may be used to find the dew-point. Put water into a polished tin and drop in pieces of ice. Stir with a thermometer and, as the ice melts, the temperature of the water inside and the air outside the tin will fall until it is low enough for a film of dew to be formed on the outside of the tin. Read the thermometer when this occurs and thus obtain a rough value of the dew-point.

Regnault's hygrometer.

Much more elaborate instruments, known as hygrometers, have been devised for making accurate determinations of the dew-point. The most accurate of these was devised by Regnault. It

consists (see Fig. 92) of a glass boiling tube, the lower end of which is replaced by a silver cup, since silver is a much better conductor than glass. The cup is cooled by pouring ether into it, and gently evaporating the ether by drawing air through it.

As soon as the dew appears the reading of the thermometer is taken and the flow of the air is stopped. The cup slowly warms up again and the temperature is again noted when the dew begins to disappear. The mean of the two temperatures is taken as the dew-point.

The observer must be very careful not to breathe on the cup during the experiment, and it is advisable to make observations from a distance through a telescope. A similar cup, not containing ether, may be placed alongside so that the surfaces can be compared; this facilitates the difficult observation of the exact moment when the dew appears and disappears.

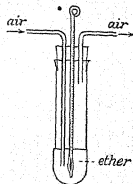


Fig. 92

Relative humidity.

Relative humidity

$$= \frac{\text{Pressure of water vapour in the atmosphere}}{\text{Saturated vapour pressure at the same temperature}}$$

Thus, if the water vapour in the atmosphere is only half saturated, the relative humidity is 0.5 or 50 per cent.

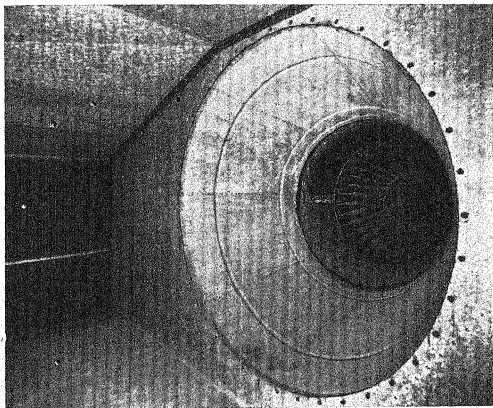
Now the dew-point is the temperature to which air must be cooled for dew to be deposited. It is therefore the temperature at which the water vapour present in the air is saturated. Hence the pressure of the water vapour in the atmosphere is equal to the saturated vapour pressure at the dew-point. Thus

Relative humidity

$$= \frac{\text{Saturated vapour pressure at dew-point}}{\text{Saturated vapour pressure at temperature of atmosphere}}$$

Hence if we find the temperature of the atmosphere and the dew-point we can, using a table of saturated vapour pressures at different temperatures, calculate the relative humidity.

vapour and the temperature at which it leaves the spray chamber is adjusted so that, on passing through a final heater, its relative humidity is reduced to the requisite value by the rise in temperature. The studios are maintained at a temperature of 70° F. and a relative humidity of 55 per cent. In all there are thirty-two fans delivering 614 tons of air per hour.



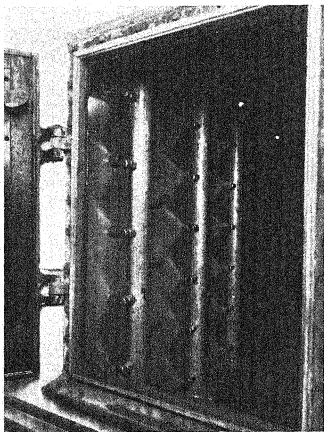
By courtesy of the B.B.C.

Fig. 93. Air-conditioning plant at Broadcasting House. An interior view of one of the large intake fans.

During the weaving of cotton, and also the manufacture of paper and cigars, a very high relative humidity of the atmosphere is essential. The humidity is carefully controlled since it is not allowed by law to exceed a certain value in order to safeguard the health of the workers.

Another commercial application of the scientific study of humidity is the artificial seasoning of timber. The moisture

content of the timber has to be reduced from 50 per cent. to about 14 per cent. in this country, when it is in equilibrium with the moisture in the atmosphere. Natural seasoning takes two or three years but the process can be carried out in a few weeks in a seasoning kiln, the temperature and humidity of which are carefully controlled.



By courtesy of the B.B.C.

Fig. 94. Air-conditioning plant at Broadcasting House. The air is passed through this water-spray, formed by pumping water through specially designed nozzles, and is thereby washed of impurities. Its humidity may be controlled by adjusting the temperature of the spray.

The wet and dry bulb hygrometer.

Regnault's hygrometer requires a skilled observer, and it is unsuitable for use in draughts. It is therefore used as a standard instrument against which other more practical, though less accurate, hygrometers are calibrated.

A form of instrument which is used in factories and meteorological observatories for finding the relative humidity is the wet and dry bulb hygrometer. It consists of two thermometers, one of which has its bulb kept damp by surrounding it with a piece of muslin or wick dipping into water (see Fig. 95).

If the water vapour in the atmosphere is not saturated, water will evaporate from the muslin surrounding the wet bulb. Since the water as it evaporates extracts its latent heat from the bulb the reading of the thermometer will be lowered. The rate at which the water evaporates and consequently the drop in temperature depends on the relative humidity of the atmosphere. By means of specially devised tables the relative humidity may be found from the reading of the dry bulb, and the depression of the reading of the wet bulb, thermometers.

The accuracy of the instrument is increased if a current of air is made to flow past the bulbs.

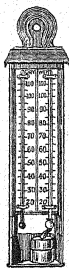


Fig. 95

The hair hygrometer.

It is well known that sea-weed becomes dry or wet according to the humidity of the atmosphere. It is said to be hygroscopic. Catgut and hair are also hygroscopic and they contract when they become wet. The former is used in a toy which is a rough type of hygrometer. Two figures are connected by a piece of twisted catgut: the "old man" appears when the air is damp and the catgut shortens, and the "old woman" when the air is dry and the catgut lengthens.

The hair hygrometer works on a similar principle. The shortening and lengthening of a hair is magnified by a suitable mechanism which moves a pointer, giving the reading of the relative humidity direct. The instrument, however, is not very reliable and needs constant recalibration.

Fog and mist.

Fog and mist are often caused by the mixing of two air currents, one of which is at a higher temperature than the other, and contains a large amount of water vapour. The lowering of the temperature of the warm air current causes it to become super-

saturated with water vapour and the vapour condenses in the form of drops.

The condensation of the water vapour in one's breath on a cold day is a simple example. The dense fogs experienced off the coast of Newfoundland are caused by the mixing of cold northerly winds from Labrador and the warm moisture-laden southerly winds from the tropical Atlantic.

Mists can sometimes be seen forming over a river at night. The temperature of the ground falls rapidly at night and cools the air above it. This cold air drains into the bottom of the river valley (just as a heavy liquid drains to the bottom of a vessel), and on coming into contact with the warmer and vapour laden air over the river cools the latter below its saturation point. Water drops condense, forming a mist.

It has been found by experiment that dust particles are essential for the formation of fog and mist. London is peculiarly liable to fog owing to its situation in a river valley and also to the smoke particles in its atmosphere.

Clouds.

Clouds, like a mist, consist of tiny drops of water. This is apparent to anyone who walks through a cloud on the top of a mountain.

When air rises it expands owing to the drop in pressure, and the work it does in expanding is taken from the heat it contains. The cooling due to this effect is 1° F. for a vertical rise of 180 ft. If the air contains a considerable amount of water vapour, this eventually reaches its saturation point and, on further cooling, drops are condensed, forming a cloud. When the air is saturated the rate of cooling with increase in height is reduced to 1° F. per 300 ft., since the water vapour in condensing gives out its latent heat.

Clouds are often formed by winds being deflected upwards by the side of a mountain. Fig. 96 shows a cloud forming on the windward side of the Aran mountains at sunrise in North Wales.

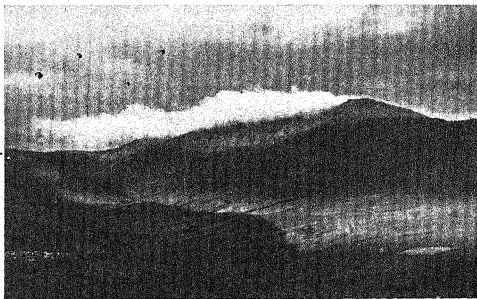
A simple experiment may be performed in the laboratory to illustrate the formation of clouds by the expansion of moist air. Take a large bottle, pour a little water into it, and fit it with a rubber bung through which passes a glass tube. Pump air into it by means of a football pump. Give the air some time to cool

down since it will have been heated owing to its compression. Now pull out the bung. The sudden expansion will cool the air sufficiently to cause a mist or cloud to form in the bottle.

Clouds (see Figs. 97–100) are classified into four main types:

1. *Cirrus*, (Latin—a curled hair)—consists of ice particles formed at heights of five or six miles.

2. *Cumulus* (Latin—a heap)—large white billowy clouds usually at a height of about one mile.



By courtesy of D. G. A. Dyson, Esq.

Fig. 96. Clouds forming on the windward side of the Aran Mountains, N. Wales. The air is forced upwards by the sides of the mountains, expands and cools: whereupon moisture condenses into tiny drops, forming clouds.

3. *Stratus* (Latin—a layer)—uniform layer of cloud, lower than cumulus.

4. *Nimbus*—dark overhanging shapeless clouds from which fall rain and snow.

Rain, hail and snow.

The water drops of which clouds are composed are extremely small, about one-thousandth of an inch in diameter, and they fall very slowly in still air—about 1 in. per sec. If, however, the

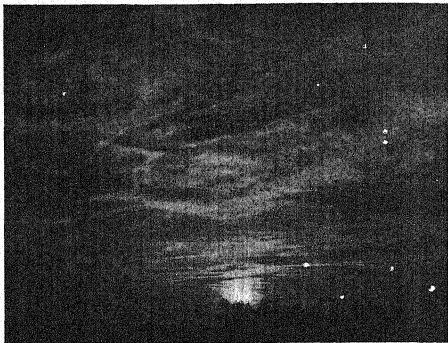


Fig. 97. Cirrus.



Fig. 98. Cumulus over South Downs.

When a relatively small area of ground is heated more than the surroundings it gives rise to a "chimney" of hot air which is cooled on ascent and so forms a cumulus cloud. Use is made of this fact in gliding. The glider aims for a cumulus cloud because he knows that there will be a rising column of air beneath it which will enable him to gain height.

clouds are forced upwards by the side of a mountain, or an upward air-current, the air is still further cooled, more vapour condenses, and the drops grow in size until eventually they are large enough to fall as rain.

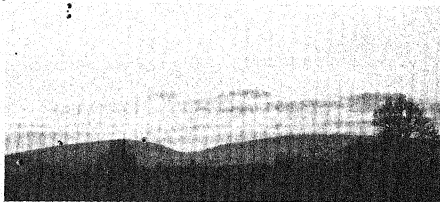


Fig. 99. Stratus over the South Downs.

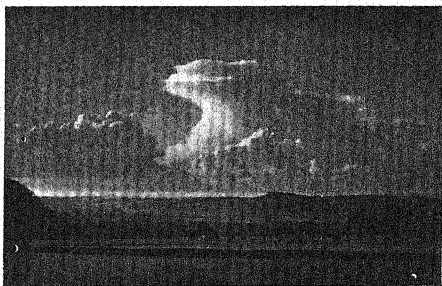


Fig. 100. Cumulo-nimbus with large false cirrus top.

Hail consists of raindrops which have been frozen by the cooling below zero of the air which has whirled them aloft to a great height.

A snowflake consists of ice-crystals which have grown on to a tiny frozen drop as it falls. You should examine freshly fallen snow with a magnifying-glass. The patterns of the crystals are very beautiful.

SUMMARY

The temperature to which air must be cooled for dew to be deposited is known as the *dew-point*.

Relative humidity

$$= \frac{\text{Pressure of water vapour in the atmosphere}}{\text{Saturated vapour pressure at the same temperature}}$$

$$= \frac{\text{Saturated vapour pressure at dew-point}}{\text{Saturated vapour pressure at temp. of atmosphere}}$$

The relative humidity may be determined by (1) Regnault's hygrometer, (2) the wet and dry bulb hygrometer, (3) the hair hygrometer.

Fog, mist, clouds, rain, hail and snow are formed by the condensation of water vapour in the atmosphere.

QUESTIONS

1. Explain what is meant by the term dew-point.

Describe how the dew-point can be determined for the atmosphere in a room.

Explain carefully what effect, if any, there will be on the dew-point (a) if a quantity of water is gradually sprinkled in the room, (b) if the temperature of the atmosphere in the room is raised. (N.)

2. Explain:

(a) Sometimes the cloud of steam from the engine of a train is long and trailing, and at other times it is very short and rapidly disappears.

(b) After a shower of rain on a sunny spring day the roads often appear to be "steaming".

(c) Following a fall of hail on a hot summer day the ground remained covered with ice particles for some minutes. Wisps of steam then appeared floating near the surface of the ground.

3. Define *dew-point* and *relative humidity*.

Explain how the values of these quantities affect the drying power of the air.

Describe a method of measuring the relative humidity of the air.

(C.)

4. Explain carefully:

(a) When a kettle is boiling the white cloud of "steam" is not visible very near to the spout.

(b) An iceberg is often surrounded by fog.

(c) On a hot day when the air is moist we feel much warmer than when the air is dry and at the same temperature.

5. It is found that on some days a deposition of moisture occurs on the inside of a window and on other days on the outside. State the conditions under which each of these deposits may occur and explain the processes involved.

(L.)

6. What is the difference between a saturated and unsaturated vapour?

Show in tabular form the results of decreasing the volume, increasing the volume, raising the temperature and lowering the temperature of a space containing saturated water vapour with some excess of water.

7. What do you understand by the expression "dew-point"?

The temperature of the earth's surface on a starlight night is observed to fall at a fairly steady rate of 5°C . per hour until the dew begins to form. Would you expect the rate of fall to be greater or less after dew begins to form? Give reasons for your answer.

(O.)

8. How is it that a thermometer whose bulb is wrapped round with damp muslin usually indicates a lower temperature than another, by its side, whose bulb is not covered? Point out how the difference of the readings might be expected to vary with external conditions.

(L.)

9. The temperature of the air is 14°C . and the dew-point is 5°C . Calculate the relative humidity. (See table of vapour pressures on p. 148.)

10. After a heavy shower of rain it is noticed that a wooden fence on which the sun is shining is "steaming". Explain the various physical processes which are involved in producing this effect.

(L.)

11. Explain how clouds are formed. Describe the four chief types of cloud forms.

12. Write a brief account of the nature and mode of formation of clouds, snow, hail and fog.

(O_a)

ANSWERS TO QUESTIONS

CHAPTER I (page 14)

6. -23.0°C .
 7. 10.6 mm.
 9. 210°F .
 13. (i) (a) 85°C . (b) 36.9°C . (c) -17.8°C .
 (ii) (a) 59°F . (b) 122°F . (c) -459.4°F .
 (iii) 75°C , 167°F .
 14. -40°C . or F .

CHAPTER II (page 33)

2. 0.16 cm.
 3. 2.2 in.
 4. 43.4°C .
 5. 30.16(4) in.
 7. $583\frac{1}{2}$ and $416\frac{2}{3}$ gm.
 9. 40.0°C ., 66.7°C .
 15. 757(0) mm.
 16. 16.0 c.c.
 18. 0.0000228 per $^{\circ}\text{C}$.
 19. 12.3 gm.
 20. 0.24 mm.
 21. 28.8 cm.
 23. 13.37 gm. per c.c.
 24. 0.000612 per $^{\circ}\text{C}$.

CHAPTER III (page 49)

2. 26 cm. of mercury, 32.5 cm.
 3. 75 cm.
 4. 76.1 cm. of mercury.
 8. 0.00369 per $^{\circ}\text{C}$., -271°C .
 9. 0.933 gm. per c.c.
 10. 171°C .
 11. $1\frac{1}{2}$ atmospheres.
 12. 91 lb. wt. persq. in.
 13. -65.8°C .
 14. 18.0°C .
 15. 689.5 mm.
 16. 783.4 mm.
 17. 21 times per min.
 18. 2.92 atmospheres.
 19. 370°C .

CHAPTER IV (page 62)

3. 55.4 gm., 0.11.
 4. (a) 0.20, (b) 0.21.
 7. $23(1)^{\circ}\text{C}$.
 8. 3.6°F .
 9. 1.3d.
 11. 21750 cal.
 12. 0.113.
 13. 439°C .
 17. 0.213, 6.04 gm.
 18. 2 : 5.
 19. 2.336×10^8 cal.
 20. 1.7°C .
 22. 0.353.
 23. 3 min. 4 sec.

CHAPTER V (page 90)

4. 30.7°C .
 5. 40°C .
 6. 41.3 gm.
 7. 11.6 gm.
 8. 299 sec.
 9. 7225 cal.
 10. 6.875 lb. of ice melted.
 13. 0.11.
 14. 25420 cal.
 15. 8°C .
 16. 120 gm.
 17. 690 gm.
 18. 0.68 lb.
 19. 32.0 lb.
 20. 10 kgm.
 21. 0.65, 9600 cal.
 22. -183°C .
 23. 8 cal. per gm.
 41. 0.9°C . too low.
 42. 55.5 cm. of mercury.

CHAPTER VI (page 104)

6. 0.0712.
 7. 0.206°F .
 8. 4.42 ergs per cal.
 9. 0.26°C .
 10. 4.27°F .
 11. 672 metres.
 12. 1167 h.p.
 13. 103.7 min.
 14. 140 min.
 15. 42400 B.Th.U.
 16. 10.8 h.p.
 17. 67760 ft.-lb.,
 87.1 B.Th.U.
 18. (a) 420 cal.,
 (b) 34300 cm. per sec.

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